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ABSTRACT

Experiments to provide an understanding of the principles related to the pressure cooker are presented. Objectives included are designed to provide the learner with the ability to calibrate a thermistor for measuring temperature; explain the meaning of latent and specific heat; calculate latent and specific heat; use a Bourdon tube pressure gauge to measure pressure; and to calculate the force on a surface. Experiences are provided so that the learner also is able to use the General Gas Laws, the phase diagram of water, and operate and explain the physical properties of pressure cookers. (SA)

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TECH PHYSICS SERIES

THE PRESSURE COOKER

A MODULE ON

THE PROPERTIES OF MATTER

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TERC Headquarters at:
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THE PRESSURE COOKER

An Introduction

WHY STUDY THE PRESSURE COOKER?

The pressure cooker which is usually found in the kitchen may hardly seem like an exciting device for study in a technical physics course. But the operation of this simple household tool is based on a number of ideas that can be quite useful to a technician. For one thing, there are some basic questions about matter -- What forms can it take? How do we describe them? What terms do we use? And there are some more practical questions relating to temperature and pressure -- What do they mean? How do we measure them? How are they controlled?

In this module you'll use the pressure cooker to explore some of these questions. Among other things, you'll learn some of the techniques that are used for measuring temperature and pressure, and about the mixed units that are used for specifying them. You'll study a bit about water and see how it behaves with changes in temperature and pressure. You'll see how important energy is in freezing and boiling and how a whole variety of systems are designed to take advantage of the energy involved in these "phase changes". At the end of this module you should have developed considerable respect for the physics of the pressure cooker and a greater appreciation for common, everyday H_2O.

A LITTLE ABOUT WHY THE PRESSURE COOKER IS USED:

We know that water ordinarily boils at 212°F, but in a household pressure cooker that can be raised to 255°F. A small amount you say, but a very important difference for many processes. For example, the process of "cooking" is a set of chemical reactions which occur in the food. Around these temperatures chemical reactions occur twice as fast for every 18° F increase in temperature. Therefore, a small increase in temperature can mean quite a shortening in cooking time. The killing of germs also depends on chemical changes that occur within them. Raising their temperature shortens their lifetime considerably and pressure cookers (often by other names) are commonly used as sterilizers.

But why use a pressure cooker and not a flame or electric heater or any of a number of other devices? The pressure cooker offers an accurate temperature-controlled environment. By controlling the pressure in the cooker (which is fairly easy to do) we can control the temperature of the boiling water. For the home pressure cooker, the pressures required are about 15 pounds per square inch. While that doesn't sound like very much pressure, it means a force on the lid of hundreds of pounds. And controlling this pressure is all done with a little three ounce weight.

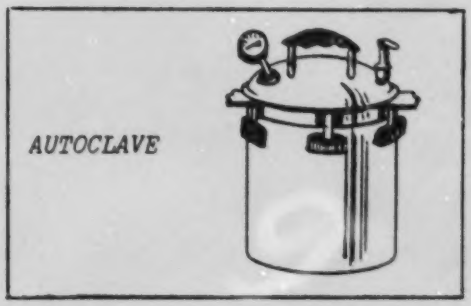
Below are a few examples of where you may encounter a pressure cooker -- or the same device by another name:

FOOD PREPARATION:

Pressure cookers are widely used for cooking food in the home and in the commercial food preparation industry. Not only can foods be cooker up to 10 times faster, but pressure cooking retains the vitamins, minerals, and natural flavors in foods better than other forms of open cooking.

STERILIZATION:

In medicine sterilization of medical instruments is an extremely important procedure. The autoclave, a form of pressure cooker,is routinely used in medical and other laboratories for fast, reliable sterilization. The common pressure cooker is often used for the same purpose to sterilize baby bottles at home.



PRESSURE MEASUREMENT:

Sometimes it is important to know the atmospheric pressure with great accuracy -- better accuracy than can be obtained with a barometer. The hypsometer is a device in which the temperature of a container of boiling water is measured with great accuracy. Changes in atmospheric pressure immediately result in changes in boiling temperature. Since the relation between atmospheric pressure and altitude is well known, such a device can also be used as an altimeter.

MODULE OBJECTIVES:

The general goals of this module are to give you an understanding of the physical principles of underlying the operation of the pressure cooker.

Among these are:

* The manner in which matter is classified into distinct phases.
* A way of mathematically describing one of these phases, the vapor phase.
* A graphical way of describing the physical conditions under which the various phases will occur.
* The effect of pressure on the temperature at which phase changes will occur.
* The energy associated with the various phases and phase changes.

In performing the module you will learn not only how the pressure cooker works but also a way of measuring temperature in remote places using a thermistor. You will learn some of the various units for measuring energy and pressure and how to convert easily from one set to another. And finally you will learn how to measure some basic weather quantities and to understand their meaning in terms of the graphical description of the properties of water.

At the end of the module you should be able to demonstrate your understanding of its objectives by doing the following:

1. Properly calibrate and use a thermistor for measuring temperature.
2. Explain the meaning of latent and specific heat and the relation between the specific heat of water and two systems of energy units.
3. Calculate latent and specific heat from sample data of the change in temperature with time of a body supplied with a known energy input.
4. Properly use a Bourdon tube pressure gauge to measure pressure, (including reading the units) above and below atmospheric pressure.
5. Calculate the force on an arbitrary surface when the pressure on that surface is given.

6. Describe the General Gas Law in terms of its parameters (including the proper units), giving qualitative examples of the relations between various sets of terms.

7. Calculate, using the General Gas Law, expected changes in one of the parameters of a gas (P, V, T) when another parameter is changed.

8. Explain, using the phase diagram of water, relative humidity and dew point.

9. Measure the dew point and relative humidity using the proper equipment, techniques and data.

10. Describe the physical principles underlying the operation of the pressure cooker.

11. Operate a pressure cooker, exercising proper precautions and procedures and explaining the physical reasons behind them.

PREREQUISITES:

Is this at your level? All you need to understand the material in this module and perform the experiments is basic laboratory skills, an introduction to scientific notation, and some algebra and graphing.

MATTER & ENERGY

INTRODUCTION

To begin our study of the pressure cooker, we must first get a bit of perspective about matter. For example, how is it classified? What are the terms used to describe its condition? What are the roles of energy and changes in energy? In this part of the module we'll explore some of these questions by a simple experiment with ordinary H_2O. We'll see that the energies required to change its condition or state are large enough so that this simple phenomenon can be used in a wide variety of common devices. In the process, we'll also learn a convenient way of measuring temperature by simply measuring the electrical resistance of a device called a thermistor.

MATTER AND ITS PHASES: The Greeks and Us. . .

Long ago man recognized that matter occurs only in a small number of different conditions, each with its own distinct properties. The early Greeks called these conditions earth, water, air and fire. Earths were substances which were solid and had a definite shape; waters flowed and would fill containers; airs could completely fill a volume yet could be easily compressed to change size; and fire was hot and could convert materials from one condition to another. To the Greeks, a tree would be earth, liquid mercury would be water, steam would be air, and so on.

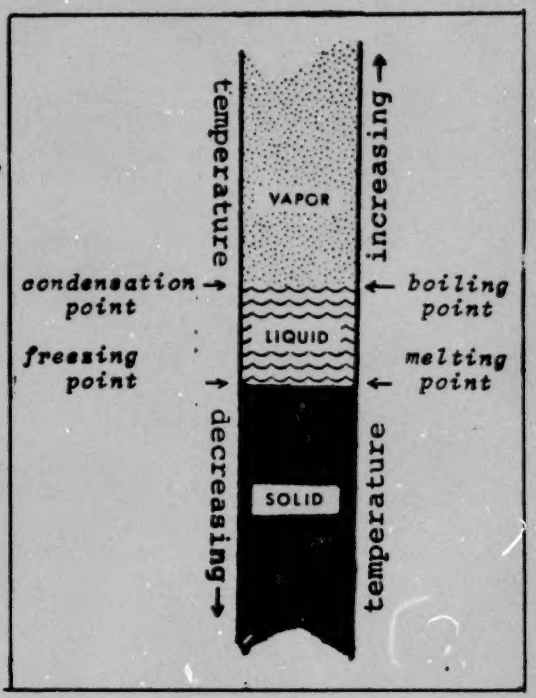
Since then, however, we've come to recognize that three of these substances, earth, water and air, were simply commonly occuring examples of a broader classification of matter. We call them the states or phases of matter: the solid phase, the liquid phase and the gas or vapor phase. But the Greeks weren't far wrong -- earth is a solid, water is a liquid and air is a gas -- and the properties the Greeks defined for the phases are largely those we use today.

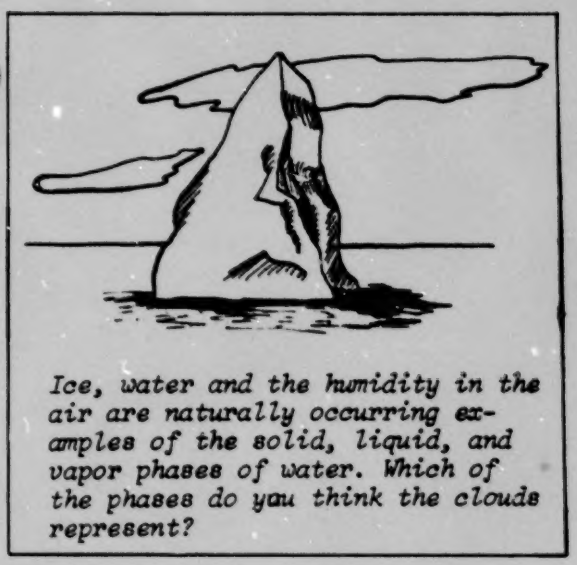
ENERGY AND PHASE CHANGES. . . .

The difference between our thinking and that of the ancient Greeks' is the idea that when a material changes its phase, for example water to ice, its H_2O molecules don't change from H_2O to become "earth" molecules, they only change their arrangement to that of a solid.

And that's where the Greeks' "fire" comes in. Fire is a source of energy and it is energy which causes materials to change phase. By adding energy we can cause a material to change from a solid to a liquid, which we call melting, or from a liquid to a vapor, vaporization. Or, by removing energy we can reverse the phase changes. These we call condensation and freezing. Of course, we don't say, "Let's add some energy to melt the ice".

As we'll see later in the module, energy, heat and temperature are all related. The point we want to make here is that melting and freezing, vaporization and condensation describe the phase changes. These phase changes occur at specific temperatures; not over a range of temperatures. Which of the four terms you use depends only on which phase change you're talking about and in which direction you're going.





Ice, water and the humidity in the air are naturally occurring examples of the solid, liquid, and vapor phases of water. Which of the phases do you think the clouds represent?

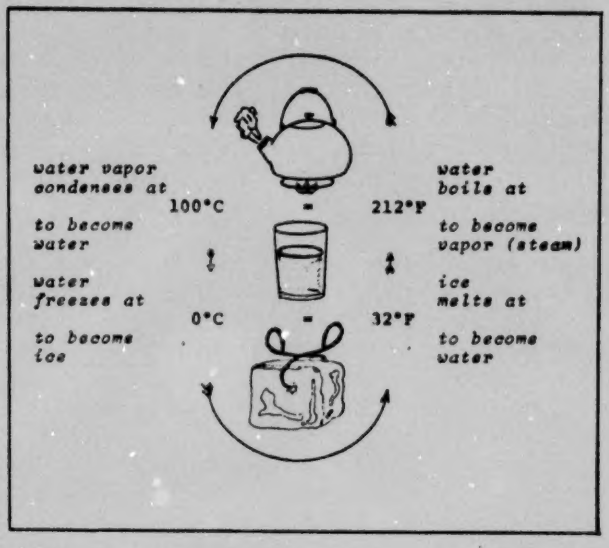
WHY IS WATER SO IMPORTANT. . . .

Perhaps you've never thought about it, but water is about the only substance that is naturally found in any significant quantity as a liquid. (Can you think of another?) Not only that, but it naturally occurs in all three of its phases. (Can you think of another one here?). We know it as a solid in the form of snow and ice, as a liquid in oceans and rivers, and as a vapor, for example, as humidity in the air. The close parallel between the range of temperatures of the liquid phase of water and the range of temperatures found on earth is one of the essential keys to the evolution of life and our environment as we know it.

It's not surprising, then, that whatever field you're in, water will probably be an essential ingredient in one phase or another. For example, water is a critical ingredient in most of the biological processes; in chemistry, water appears in untold reactions and processes; one of the essential concerns of ecology is pure water; oceanography would be a dull subject without water; in meteorology all three phases of water are of critical importance; since 65% of your body is water it can't help but be essential in medicine; and so on.

THE PHASES OF WATER. . . .

The three phases of water and the temperature at which it changes phase are well known to us. Water boils at 212°F (or 100°C), changing from a liquid to a vapor; and it freezes at 32°F (or 0°C) to become a solid. The phases are so common to us that we have special names for them. We call the vapor phase steam and the solid phase ice. For no other substance do we do that, and the reason is that we seldom encounter other materials in more than one of their phases.



. . .AND OF OTHER SUBSTANCES:

Since we rarely see them, do other substances occur in all the three different phases?

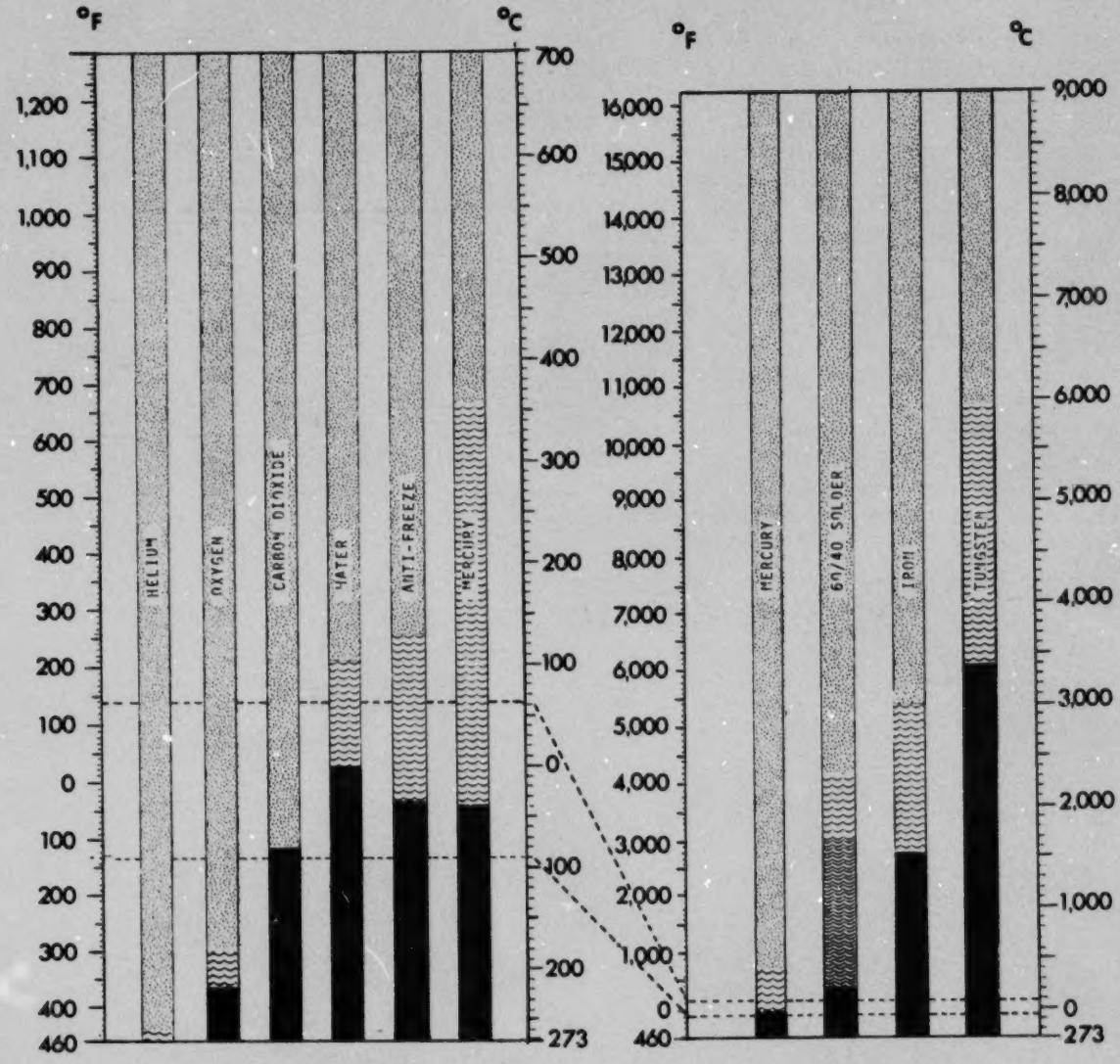
The Usual Case: In general, yes, most substances can occur as a liquid, vapor, and solid, but it often takes special conditions to produce them. We all know solid iron, for example, but few of us have seen liquid iron or gaseous iron. These phases can occur, however, but only at very high temperatures. Iron becomes a liquid at 1808°F and vaporizes at 3273°F. We all know that oxygen is a gas, it is an essential part of the air we breathe. However, oxygen does not become liquid until the temperature falls to -361°F, and becomes solid at -424°F. Mercury is a liquid at normal temperature, but it becomes a gas at 630°F. Most of these temperatures are far beyond those that occur naturally on earth.

And Exceptions: Usually, as the temperature increases, a solid changes first to a liquid, then to a gas. But, there are some interesting exceptions: carbon dioxide changes directly from a solid to the gas phase. Have you ever seen dry ice? This is solid CO_2 which converts directly to a vapor. Such a phenomenon is called sublimation.

Another exception is substances which do not have definite melting points or boiling points. A familiar example is glass: as it is heated, it does not suddenly change from solid to a liquid at a definite temperature; rather, it becomes softer and softer until finally it flows quite easily as a liquid.

AT ONE ATMOSPHERE!

There is one extremely important qualification we must give to the boiling and freezing temperatures and to the descriptions we have given on these pages. This is that *these temperatures only apply when the fluid is at atmospheric pressure.* When the pressure on the fluid changes, these temperatures change and that, as we mentioned earlier, is the principle behind the functioning of the pressure cooker. We will study these changes quantitatively for water in Section 3.



In the diagram above, the range of temperature for the solid, liquid, and vapor phases of a number of different substances are given. Note that the left-hand group is on an expanded scale compared to the right-hand group. The dotted lines indicate the highest and lowest temperatures ever recorded on earth. Some interesting points to note are: helium is the only substance that does not become a solid at one atmosphere; carbon dioxide is never a liquid at one atmosphere but sublimes directly from a solid to a vapor; the data for anti-freeze is for a 50/50 mixture of anitfreeze (ethylene glycol) and water; 60/40 solder is an alloy of tin and lead that has a single freezing point but the constituents vaporize separately at their own boiling temperature.*

* *Highest recorded temperature:* 134°F, Death Valley, 1913
Lowest recorded temperature: -126.9°F, Anarctica, 1960

A SIMPLE EXPERIMENT WITH H_2O

Having classified matter into three phases and discussed the phase changes we can now investigate a bit what it is that determines which phase a particular material is in. As in most of physics, the key quantity is energy and in the experiment we will explore how energy is involved in changing phases of matter. Since we would like to look at all three phases of matter we want to choose a material for which all three can be studied. And this of course, leads us back to water.

WHAT WE WANT TO DO:

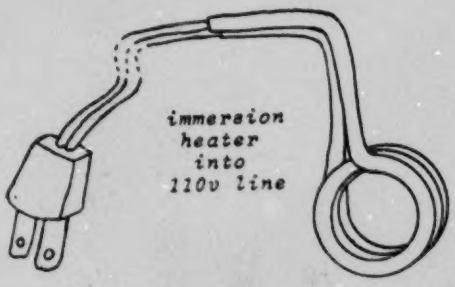
SUPPLY ENERGY AT A CONSTANT RATE to A KNOWN MASS OF H_2O and

You must be able to put energy into the water at a constant rate so that you can measure the energy input accurately. You can do this easily with an electric heater: You can control the rate of energy input (the power) and measure it accurately with simple instruments.

If you are to measure something useful about water, you need to know the amount of water you started with; this amount is best described by measuring the <u>mass</u> of water.

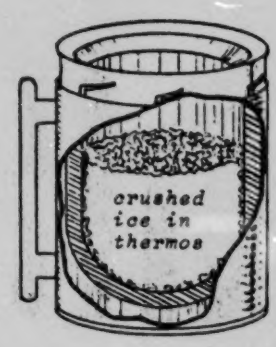
WHAT WE WILL USE TO DO IT:

IMMERSION HEATER



A simple immersion heater, that you might use to heat water for coffee, serves well for our purpose. When the heater is connected to an electrical outlet, the heater "heats up" and its temperature rises. There is heat flow from the heater to the water. What began as electrical energy becomes internal energy of the water and the heater.

FROZEN MASS OF ICE



You will work with water in a Thermos bottle to cut down heat losses. One way is to freeze the water with the temperature sensor and the heater in it. If a freezer is not available, then you can start with a mass of crushed ice.

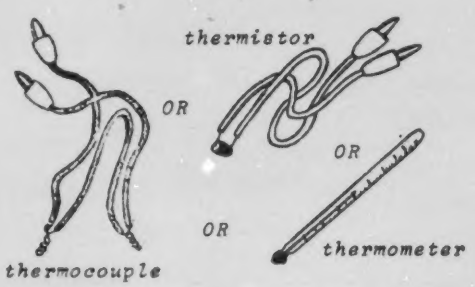
By starting with water in its solid phase and adding energy at a constant rate we can find how much energy it takes to raise its temperature and to cause phase changes. We'll see the relation between energy supplied and temperature. And we'll find that just as the properties of water (boiling and freezing point) were used to define the two common temperature units, another set of properties are used to define two common energy scales. Finally we'll see how the large amount of energy required to cause a phase change is used in countless ways in technology. The discussion below and on the next page will provide a simple analysis of how one goes about setting up an experiment like this to get the data that is needed.

MEASURE THE CHANGE IN TEMPERATURE with TIME

The heart of this experiment is measuring the temperature of the water as time goes by. By knowing the temperatures at definite times, you can determine the response of the water to a definite input of energy.

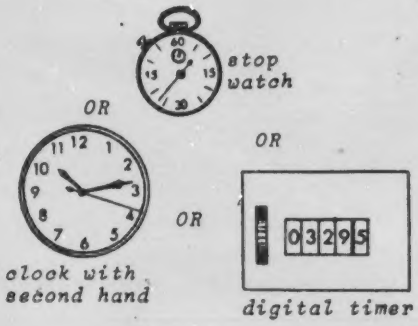
Time is the essence: Energy is flowing into the water at a definite rate. The time elapsed since turning on the electrical power determines the total amount of energy put into the water.

THERMISTOR



A thermistor is a good "thermometer" for this experiment. They are inexpensive, easily available, and widely used in technology. Also, they are small and therefore ordinarily have little effect on the temperature being measured. For the same reason, they respond quickly to changes in temperature. An ordinary thermometer could be used, but its large mass would have a significant effect on the water temperature and its response would be slow. A thermocouple would require more elaborate apparatus.

TIMER



A convenient timer would be a stop watch or a Kodak Darkroom Timer. They can be set to read zero at the start. An electric clock with a clearly marked face and a sweep second hand also serves well. Do not use an electric clock that has only 4 divisions (3 marks) between the hours.

DATA WE'LL NEED:

$\frac{\Delta Q}{\Delta t}$ = RATE OF ENERGY IN ------

You'll need to know the rate of energy being supplied to the water by the immersion heater. This rate (called the power) is the energy flowing in per unit of time. When energy, ΔQ, is measured in *joules*, the rate is measured in joules per second. $\Delta Q/\Delta t$ Joules per second are simply the power input in watts:

$\frac{\Delta Q}{\Delta t}$ is in watts = joules/second

The power supplied to the water by this immersion heater is generated electrically so that a calculation of the *electrical* power produced by the immersion heater will give us the rate of energy in. The immersion heater will be plugged into an ordinary 110V ac outlet and the equation for ac power generated in such a resistive heater is:

$$P = \frac{V^2}{R} \text{ which equals } \frac{\Delta Q}{\Delta t}$$

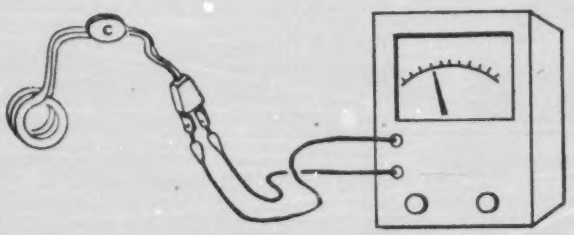
where V is the ac, rms* voltage supplied in volts and R is the resistance of the heater in ohms.

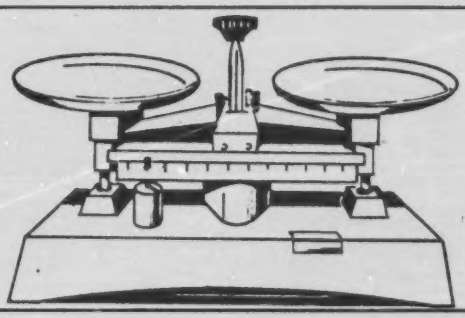
M = MASS OF H_2O ------

You must know how much H_2O you have which means knowing its mass. We use the word *mass* to describe the amount of water. This is different from the more familiar term *weight* which has a different meaning. Weight is a *force*, the force of gravity pulling on a mass. Mass and weight are of course closedly related; the greater the mass of an object the greater the weight. A balance shows when the force of gravity is the same on two objects, and thus indicates when their masses are equal.

One metric unit of mass is the gram. The gram is a small unit about equal to the mass of a cube of water one cm. on a side. About 250 grams of solid or crushed ice is about the right amount for this experiment. It isn't necessary to use some exact amount but you must know the amount exactly since later calculations will use this figure. A common ice cube has a mass of about 25 gms. Therefore, the 250 gms. of ice we suggest you use is about 10 ice cubes.

HOW TO GET IT:





To calculate the power supplied by the immersion heater we need to know its resistance, and the ac, rms voltage supplied to it. The simplest way to find the resistance is to use an instrument called an ohmmeter. (See the discussion under thermistor). The resistance of the heater changes with temperature. However, most immersion heaters are made of materials with a low temperature coefficient of resistance. This means that its resistance changes only slightly with temperature. Thus despite the fact that the heater gets quite hot you can probably use the value you get at room temperature without too much error. You might check this statement to be sure.

The voltage at the outlet in the wall is between 110 and 120 volts and this figure is the ac, rms value. For a more detailed discussion of alternating current electricity check a reference book.

You will have to measure the mass of H_2O twice; once at the start and again at the end to see how much water you boiled away. For this purpose a simple balance is suitable. A good procedure is to first measure the mass of the thermos bottle container (and record the value for later referal). When the thermos has been thus balanced, add about 250 gms. of weight to the weight pan. Then add enough crushed ice until the two pans are back in balance.

NOTE: *Be sure you only put ice in the thermos and no water. An essential part of this experiment is the amount of energy it takes to convert ice to water. Therefore, having water in the thermos to start will cause errors. To minimize melting after you add the ice it is wise to pre-cool the thermos with ice before the initial weighing.*

R = THERMISTOR RESISTANCE TO GET THE TEMPERATURE--

A thermistor is a conductor of electrical current whose resistance varies greatly (but always in the same way) with temperature. Therefore, you must carefully measure the electrical resistance of the thermistor as the temperature of the H_2O changes. Thermistors have a high *negative temperature coefficient of resistance* which simply means that as the temperature goes up the resistance goes down and vice versa. We will see later in the module how to calibrate the thermistor to get the temperature but for the moment we will be primarily concerned with getting accurate resistance values.

Recall that the electrical resistance, R, in ohms means the voltage, V, across the element to produce a given current, I.

$$R = \frac{V}{I}$$

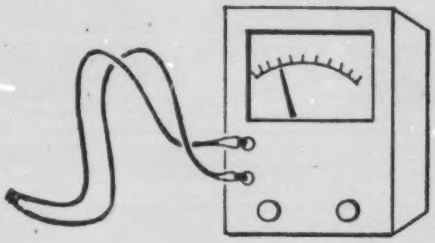
Thus applying a voltage to the thermistor and measuring the current that flows will give you the resistance.

Δt = ELAPSED TIME.....

You need an accurate record of the elapsed time because the *total* energy, ΔQ, supplied to the H_2O is proportioned to the time, Δt, it was flowing. That is when energy flow is maintained at a constant *rate*, multiplying this rate, $\Delta Q/\Delta t$, by the elapsed time gives the total energy:

$$\Delta Q = \frac{\Delta Q}{\Delta t} \times \Delta t$$

If $\Delta Q/\Delta t$ is in watts (joules/second) and Δt is in seconds, the energy will be in joules.



The simplest way to measure the resistance of the thermistor is to use an instrument called an ohmmeter. This applies a known voltage (usually from a small battery) across the thermistor and then registers the resulting current on a meter scale calibrated in resistance. Generally the ohmmeter has several ranges of resistance. You will have to switch from range to range during the measurement since the thermistor resistance changes greatly with temperature.

The only caution is that the power generated in the thermistor now cause it to heat itself (self-heating effect) and read the wrong temperature. This possibility is described by the *dissipated* constant which is the amount of the thermistor power necessary to raise its temperature 1°C. A typical value is 4m watt/°C so that you must keep the measuring power below about 1m watt. Most commercial ohmmeters will meet this requirement.

You will want to know the elapsed time in seconds since watts are joules/second. Thus if your timer measures minutes you will have to make the conversion.

This can either be done later in analyzing your data or initially when you make your data table. It usually is convenient to take your data in some easy multiples of a minute, for example 10 sec., 15 sec., or 30 sec.

THE STEPS

PREPARING FOR THE MEASUREMENT

1. Prepare a data table as shown below. In the column for notes you should record what you observe. For example note when all the ice has melted, when the water starts to boil, etc. Make several additional columns for later calculations.

2. Measure the resistance of the immersion heater; check that this value still applies when the heater is hot. Caution: *Do not plug the heater in unless it is immersed in water since it will melt. Either immerse it in boiling water or heat it lightly in a flame.* Set the voltage on a variable voltage supply to produce a heater power of about 160 watts. Record your values in the table at the right.

3. Weigh the mass of ice you start with. Be sure to pre-cool the container and minimize the amount of water.

4. Connect the thermistor leads to the ohmmeter. At the beginning it should indicate several thousand ohms which corresponds to room temperature. Record these values.

5. Insert the heater and thermistor deep in the ice, making sure they are well separated but surrounded by ice.

Heater Resistance:_____

Heater Voltage:_____

Power Input:

$\frac{\Delta Q}{\Delta t}$ = _____

Mass of Thermos:_____

Mass of Thermos & Ice:_____

_____

Initial Mass of H_2O:

M_i = _____

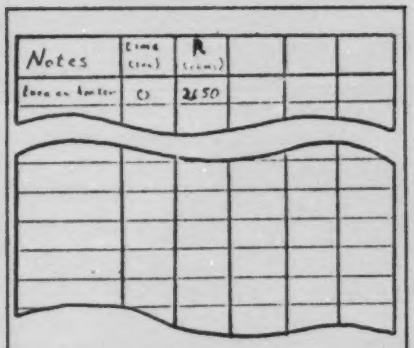
Initial Thermistor Resistance:

R = _____

Room Temperature:

T = _____

NOTE ON MAKING THE RUN



SAMPLE DATA TABLE
Showing Typical Data

6. Begin the experiment: Turn on the heater, record the initial thermistor resistance, and start the stop watch (or record the initial time).

As soon as enough ice has melted, stir the water frequently. This ensures that the temperature is reasonably uniform throughout the water, and helps maintain good contact between heater, water and thermistor.

When the temperature is changing rapidly, as when you are heating the liquid water, you should take a reading about every minute. When the temperature remains nearly constant, intervals of three or four minutes are short enough.

Be sure to note on your table of data the exact times when all the ice is melted and when boiling begins.

Boil away enough water so you can measure accurately the mass boiled away. At least ten minutes of boiling is required with the power input and mass of water you are using.

SOME FINAL DATA

7. When enough water has boiled away, turn off the heater power. Weigh the mass of water remaining and record it.

As a further check on the operation of the thermistor, try some additional measurements: take your own body temperature by carefully putting the thermistor in your mouth under your tongue. Don't short the leads and don't eat it. Check a number of other temperatures in the room, for example cold and hot tap water, drinking fountain water, a radiator, etc. Record the thermistor resistance in each case.

Final Mass Thermos & Water:___

_____

Water Boiled Away:

ΔM = _____

Body Temperature:

R = _____

Other:_____

_____

_____

_____

ANALYSIS OF THE DATA -or- What To Do With All The Numbers

Now that you have taken your data the next step is to display it in such a way that you can easily see what was going on. A graph of the temperature of the H_2O with time will do that for this experiment. From the graph we will be able to analyze what was happening at each stage and extract quantitative data of the amount of heat required to carry the H_2O from phase to phase. In order to draw the graph we must first determine the temperature which corresponds to the various resistance values of the thermistor, a process called *calibration*.

OPTIONAL

THE THERMISTOR - A Little About How It Works:

On the next two pages we will describe the calibration procedure for the thermistor. But first we want to talk briefly about how the thermistor works as a thermometer and why only a couple of reference temperatures can give you the whole calibration range.

A thermistor is a ceramic material made by sintering mixtures of metallic oxides such as manganese, nickel, cobalt, copper, iron and uranium. These materials are selected because of their semi-conducting property of a large negative temperature coefficient of resistance. This means that as the temperature goes up, the resistance goes down, just as you observed in your experiment.

The relation between resistance and temperature for thermistors can be described quite accurately by the mathematical equation:

$$\frac{R(T)}{R(To)} = e^{\beta\left(\frac{1}{T} - \frac{1}{To}\right)}$$ where:

R(T) is the resistance at absolute temperature T
R(To) is the resistance at absolute temperature To
e is the constant 2.718
β is a constant which depends on the thermistor material (typical value = 4000°absolute)

The reason why this relation describes the thermistor behavior can be explained by the theory of semi-conductors, though we shall not go into that here. However, by the following mathematical manipulation we can derive from the equation a simple relation between R and T that will let us calibrate the thermistor.

Taking the natural logarithm of both sides: $$\ln\frac{R(T)}{R(To)} = \beta\left(\frac{1}{T} - \frac{1}{To}\right)$$

and separating the logs: $$\ln R(T) - \ln R(To) = \beta\frac{1}{T} - \beta\frac{1}{To}$$

. . . . we get: $$\ln R(T) = \beta\left(\frac{1}{T}\right) + C$$

where C is a constant equal to: $$\left[\ln R(To) - \beta\left(\frac{1}{To}\right)\right]$$

Recalling the natural log-common log conversion ($\log R = .434 \ln R$) we get:

$$\boxed{\log R(T) = .434\,\beta\left(\frac{1}{T}\right) + C'} \quad \text{where} \quad C' = .434C$$

What this says is that *if we plot log R against* $\frac{1}{T}$ *we should get a straight line.* Since two points are all that are needed to determine a straight line, we need but two pairs of R, T values to draw the calibration graph. These of course we have, namely the melting and boiling point of water, plus a couple of others for good measure.

CALIBRATING THE THERMISTOR

INTRODUCTION

The expression derived on the previous page tells us that for a thermistor, log R is proportional to $\frac{1}{T}$ where R is the thermistor resistance in ohms and T is its *absolute* temperature. One really only needs two points to determine the calibration curve but since small errors do sometimes occur in measurement we will have you use four reference temperatures; freezing and boiling points of water, your mouth temperature and room temperature. Using your values for these four temperatures, follow the procedure below to get your calibration graph.

1. Convert from °C to °K by adding 273°C:

   The temperature you use must be *absolute* temperature (°K). You will learn more about absolute temperature scales later in the module but for the moment simply add 273°C to each of your temperature readings. This will give you degrees in Kelvin (°K), an absolute temperature appropriate to our measurement.

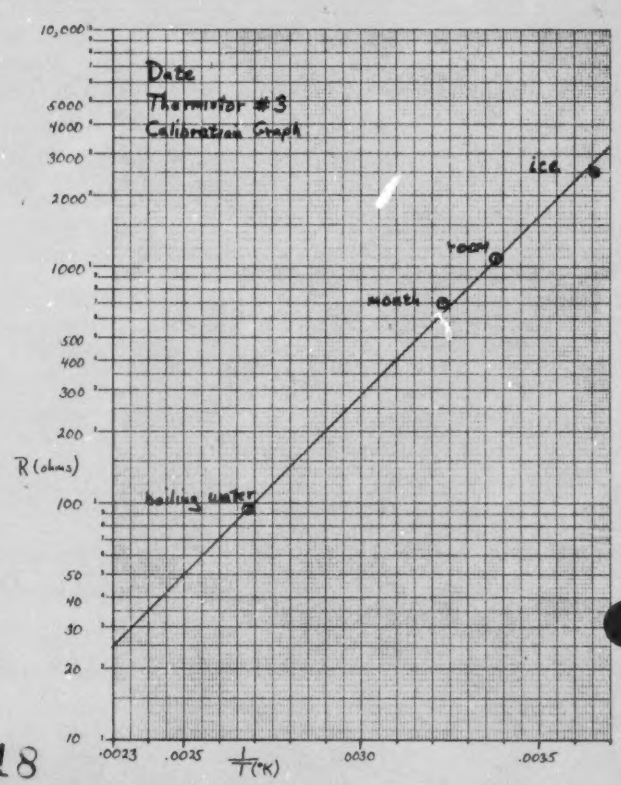
2. Calculate $\frac{1}{T(°K)}$:

   Since log R is proportional to $\frac{1}{T}$ you must calculate $\frac{1}{T(°K)}$ for each of your four values. The sample table opposite shows typical values of the data.

Reference	T (°C)	T (°K)	$\frac{1}{T}\left(\frac{1}{°K}\right)$
Melting ice	0	273	.00366
Room			
Mouth			
Boiling Water			

3. Plot R vs $\frac{1}{T}$ on Semi-log Paper:

   Rather than looking up the values of log R and plotting them for each of the resistance values it is much more convenient to use semi-log paper where the work has all been done for you. Semi-log paper has logarithmically spaced lines on the y-axis but linearly spaced lines on the x-axis. For your plotting you will need at least 3 cycles of log as shown opposite. Label your graph paper and plot your four values.



4. Draw a straight line through your points:

   Since the mathematical expression says that these four points should lie in a straight line you should try to draw the best straight line you can through the four points. In general this means that the number of points that lie above the line (and their distance) should equal the number that lie below.

The graph you have just made is now perfectly adequate as a calibration graph for your thermistor. However, the temperature scale is in $\frac{1}{T(°K)}$ which is not a terribly convenient unit. Consequently a more convenient graph can be derived from this with the x-axis marked in T(°C). Follow the procedure below to obtain this graph.

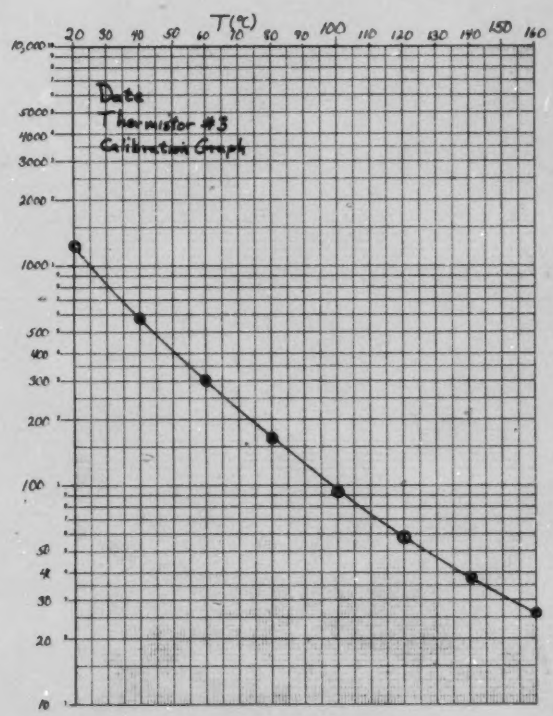
5. Determine values of R for intervals of T(°C) from your graph:

To obtain the graph you will need considerably more than four points since it will not be a straight line and more points will give you a more accurate indication of its true shape. First make a table of equally spaced values of T, say every 20°C, convert these to °K by adding 273°C, and then calculate $\frac{1}{T(°K)}$. From these values look up the corresponding value of R from your first graph. The sample table at the right should serve as a guide.

T(°C)	T(°K)	$\frac{1}{T}$ $(\frac{1}{°K})$	R (ohms)
0	273	.0366	2800
20	293	.0341	
40	313		
60			

6. Plot R vs T on Semi-log Paper:

Now plot your values of R and T, again using semi-log paper. You can probably use the same piece you used before by simply placing the T(°C) axis at the top. We use a semi-log axis for R because it covers such a wide range of values.



7. Draw a smooth curve through your points:

Connect your points with a smooth curve. A draftman's French curve is a useful tool to help draw a smooth accurate line. The more accurate your curve of course the more accurate will be you calibration data.

You now have a quite convenient and reliable calibration graph for your thermistor. This graph is only accurate for the particular thermistor you used to obtain it *so you should be sure you have noted which thermistor it is that you used.* One of the attractive features of thermistors is that their characteristics do not change appreciably with time so that your graph can be used over and over again. HOWEVER, you should always check at least one point on the graph each time you use the thermistor to assure yourself that you have the right calibration graph and that the thermistor has not been damaged.

PLOTTING THE DATA:

With your thermistor accurately calibrated, you are now able to convert your thermistor readings to temperature. Then you can graph the change in H_2O temperature with time. The procedure below can serve as a guide for producing your graph.

1. CONVERT THERMISTOR RESISTANCE TO TEMPERATURE:

   Using your calibration graph, determine the temperature corresponding to each thermistor reading of your data. Record these values in your data table.

2. PREPARE A GRAPH:

   For this experiment we want to plot the temperature of the H_2O with time. The most convenient way to do this is to put temperature on the y-axis. The range of temperature will be from 0-100°C and time, 0 to about 30 minutes. Select your divisions carefully and label your axes clearly.

3. PLOT T vs t:

   Now plot your data points. Be sure to include the notes that you took, indicating exactly when all the ice melted and when boiling began.

4. DRAW A SMOOTH CURVE THROUGH YOUR POINTS:

   Draw the best smooth curve through your points that you can. In some regions the data will fall in a straight line and a straight edge will be useful. In other regions you will have to sketch carefully. A guide to the best line is that the number of points that lie above the line (and their distance) equals the number that lie below.

YOUR GRAPH:

What is the graph that you have now drawn? To summarize it is *a graphical representation of the change in temperature of a mass of ice when energy is supplied to it at a constant rate.* Experimentally you observed that the ice first melted to water, the water then heated up and finally began to boil. As a temperature graph this becomes a flat region, an upward slope, and finally another flat region.

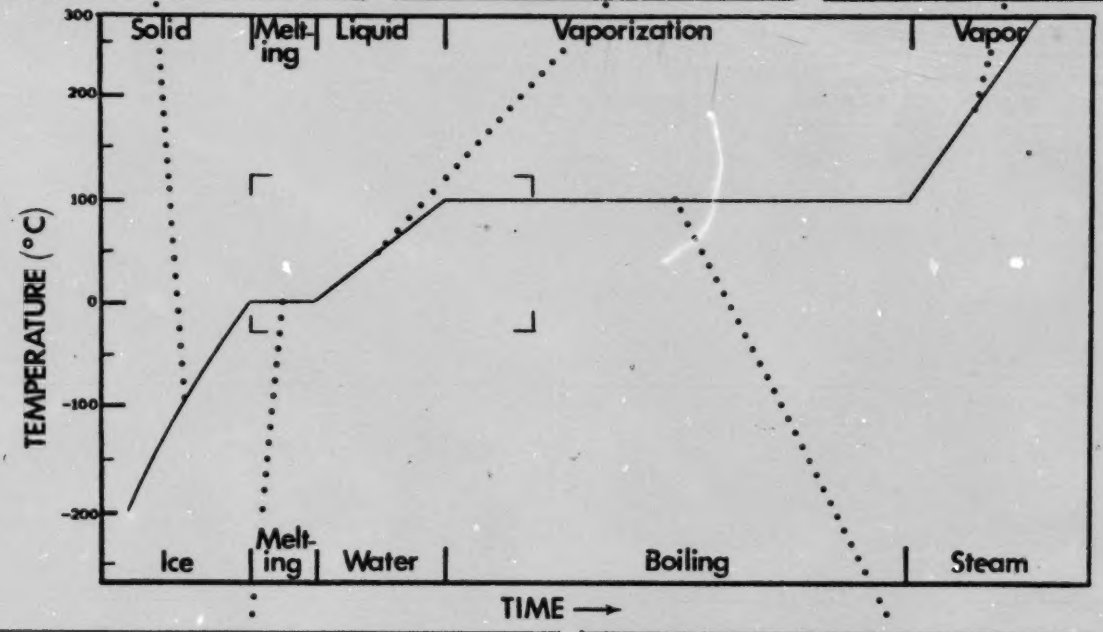
On the opposite page is a graph of an idealized experiment on water over a considerably wider range of temperatures. It illustrates what would have happened if you had started with ice that was extremely cold and continued supplying energy until all of the water converted to steam and you continued to heat the steam. To perform such an experiment would require rather sophisticated apparatus and techniques. Consequently your simple experiment covered only the region outlined in brackets and almost certainly did not yield results that were as crisp and neat.

If you could have performed a similar idealized experiment on other substances you would have found a graph that looked much like the one shown. There would be a difference, of course, in scale, length of lines and slopes but the general character of two flat regions separating three upward sloping lines would be the same. Further, the description of what was happening would have been much like that given.

If the initial temperature of the ice had been colder than 0°C, say -200°C, the input energy from the heater would have initially gone into heating up the ice. Not until the ice temperature reached 0°C would it begin to melt. Note that the line is slightly steeper at lower temperatures than it is near 0°C. This indicates that the amount of energy needed to raise the ice temperature by one degree is less at colder temperatures than it is near 0°C.

Once all the ice has melted the water starts to heat up. The energy output of the electrical heater goes into internal energy of the water and its temperature rises. Note that the line here is extremely straight, meaning that the energy required to change the temperature of the water by one degree is nearly constant. Also, the slope of the line is less steep than that for ice or steam. Thus it requires more energy to raise the energy of a mass of water by one degree than it does for an equal mass of ice or steam.

If all of the steam that came off when the water boiled away had been contained and kept in contact with the heater, its temperature would begin to rise as soon as the last drop of water disappeared. The slope of the rise is fairly constant and almost identical to that of ice near 0°C.



Once melting begins, the temperature of the ice/water mixture remains extremely constant at 0°C. All of the output energy of the electric heater goes directly into the phase change of melting with no temperature rise. Thus the liquid state must be one of higher internal energy than that of the solid, even when they are at the same temperature. This energy will be liberated if the water is made to freeze on some cold surface.

When the temperature of the water reaches 100°C its temperature stops rising and and boiling begins. All of the energy being put into the water goes directly into the phase change of vaporization. Note that the time it takes to vaporize all of the water is considerably longer than that required to melt the ice. This means that the energy required to vaporize a given mass of water is considerably greater than that required to melt an equal mass of ice.

ANALYZING YOUR EXPERIMENTAL RESULTS:

By looking at your graph of change in temperature with time you have gained only a *qualitative* understanding of your experimental results. It is more significant to determine something *quantitative*. In this case that means establishing a definition that can be used for any substance and then applying that definition to your experimental data on water. In reviewing the sample data on the previous page we see that there are two distinct types of process, heating and a phase change. The energies required to produce these results are similarly expressed by two distinct terms, specific heat and latent heat.

The specific heat is the amount of energy required to heat a unit mass of a substance by one degree.

The latent heat is the amount of energy required to convert a unit mass of a substance from one phase to another.

As usual there are different units that can be used for energy, mass and temperature so that the above definitions are quite generally stated.
If you need to refresh your memory as to the meaning of energy and power and the units used to measure them, refer to the summary pages on this topic at the end of this section.

Phase OR Phase Change	The Term	The Symbol	The Definition
Solid (Ice)	Specific Heat Of The Solid (Ice)	$C_s = \frac{Q}{M_s \Delta T}$	The amount of energy required to raise one gram of ice one degree Celsius: Q = Amount of energy required in calories. M_s = Mass of ice in grams. ΔT = Total change in temperature in degrees Celsius.
Melting (Fusion)	Latent Heat Of Melting (Fusion)	$L_m = \frac{Q}{M_s}$	The amount of energy required to convert one gram of ice into water: Q = Amount of energy required in calories. M_s = Mass of ice converted in grams.
Liquid (Water)	Specific Heat Of The Liquid (Water)	$C_l = \frac{Q}{M_l \Delta T}$	The amount of energy required to raise one gram of water one degree Celsius: Q = Amount of energy required in calories. M_l = Mass of water in grams. ΔT = Total change in temperature in degrees Celsius.
Vaporization (Condensation)	Latent Heat Of Vaporization (Condensation)	$L_v = \frac{Q}{M_l}$	The amount of energy required to convert one gram of water into steam: Q = Amount of energy required in calories. M_l = Mass of water in grams.
Vapor (Steam)	Specific Heat Of The Vapor (Steam)	$C_{vapor} = \frac{Q}{M_v \Delta T}$	The amount of energy required to raise one gram of steam one degree Celsius: Q = Amount of energy required in calories. M_v = Mass of steam in grams. ΔT = Total change in temperature in degrees Celsius.

A couple of things about these definitions should be noted. First the latent heat definition does not include temperature. This, of course, is because there is no temperature change during a phase change. Second, the definitions work equally well in both directions. That is if the substance is cooling rather than being heated, the energy is given off rather than having to be supplied. Third, if we substitute "gram of water" (pound of water) for "unit mass of a substance" and add "Celsius" (Fahrenheit) to "degree" we have the *definitions* for the calorie and the BTU respectively. Thus we see that the specific heat of water defines two of the common energy units. The table below shows the relationship between the several heats for water, shows how they may be calculated from your data and gives the accepted values.

5. CALCULATE C_l, L_m and L_v FROM YOUR DATA:

Express your results in metric (Calories) and in English (BTU's) units. Recall that one pound equals 454 grams.

Accepted Values	Your Values	Sample Calculation of C_l
.50 $\frac{\text{calories}}{\text{gm}\ ^\circ\text{C}}$ @ -2°C	X	In making a calculation of the specific heat of water one must be careful to use only the *linear* portion of the graph. Departures from linearity in this case represent a non-uniform heating of the water. For example the departure as the water approached boiling was due to the fact that there was really a bit of boiling going on around the heater.
79.7 $\frac{\text{calories}}{\text{gm}}$ 143 $\frac{\text{BTU}}{\text{lb}}$	L_m	Suppose the graph is quite linear from: t_1= 11.0min to t_2= 19.0min for which: T_1= 61°C and T_2= 12°C Then for this portion it took: Δt_h= 19.0-11.0 = 8.0min = 480sec to raise: M_l= 250 gm of water by: ΔT= (61°C - 12°C) = 49°C
Definition 1.00 $\frac{\text{calories}}{\text{gm}\ ^\circ\text{C}}$ 1.00 $\frac{\text{BTU}}{\text{lb}\ ^\circ\text{F}}$	C_l	If the rate of energy supplied was; $\frac{\Delta Q}{\Delta t_h}$ = 170 $\frac{\text{joules}}{\text{sec}}$ . . then the total energy required was; Q= 170$\frac{\text{joules}}{\text{sec}}$ X 480sec ...or converting to calories: = 82,000 joules x .24 $\frac{\text{calorie}}{\text{joule}}$ = 20,000 calories
569 $\frac{\text{calories}}{\text{gm}}$ 1074 $\frac{\text{BTU}}{\text{lb}}$	L_v	The basic equation for specific heat is; $C_l = \frac{Q}{M_l \Delta T}$
.48 $\frac{\text{calories}}{\text{gm}\ ^\circ\text{C}}$ @ 110°C	X	...which gives; = $\frac{20{,}000 \text{ calories}}{250 \text{ gm} \times 49^\circ\text{C}}$ *the specific heat of water:* = 1.6 $\frac{\text{calories}}{\text{gm}\ ^\circ\text{C}}$

COMPARING THE VALUES....

Let's now take a look at the values for specific and latent heat that you have obtained, Several types of analyses can be made. First you can compare your values to the accepted values to see how good an experiment you performed and possible causes for error. Second you can compare the relative sizes of the heats to see what physical significance they may have. And third you can compare the values for water with those of other substances. On the next few pages some of these comparisons are discussed.

EVALUATING YOUR EXPERIMENT. . .

In our sample calculation of C_l we obtained a value of 1.6 $\frac{\text{calories}}{\text{gm } ^\circ\text{C}}$. The value by definition however is 1.00 $\frac{\text{calories}}{\text{gm } ^\circ\text{C}}$. Our value was too high which says that it took more energy to raise the temperature of one gram of water by one °C then it should have. Chances are your value also was too large. Can you think of possible reasons why this was so? Was it the nature of the water you had or was it due to the apparatus you used? Below are examples of questions that you might ask yourself about your experiment to help you evaluate your values of C_l, L_m and L_v.

* Did the value differ significantly from what you might expect, and in what direction (too high, too low)?
* What were possible sources of error; for example, flaws in the design of the apparatus, malfunctions of the apparatus, inaccuracies in measuring the data, etc.?
* Can you estimate the relative importance of these error sources?
* What steps could you take to improve the experimental apparatus?
  . . . . . to improve the experimental procedure?
* Why do you think we did not suggest them?

COMPARING THE MAGNITUDES OF C_s, L_m, C_l, L_v and C_{vapor}. . .

THE SPECIFIC HEATS: In comparing the three specific heats, the first thing one notices is that the values for the three phases of water are all about the same size, all being about 1 calorie/gm °C . Thus despite the fact that the properties of ice, water and steam are quite different, the amount of energy required to raise the temperature of a unit mass remains about the same. What

does this mean? One interpretation is that there may be some fundamental relationship between energy change and temperature change, that is independent of phase. While this is hardly a proof of that assertion, other evidence does show this to be true. In the next section of the module you will see what this relationship is for substances in the gas phase.

Ice	$C_s = .50 \frac{cal.}{gm\ ^\circ C}$
Water	$C_l = 1.00 \frac{cal.}{gm\ ^\circ C}$
Steam	$C_{vapor} = .48 \frac{cal.}{gm\ ^\circ C}$

THE LATENT HEATS: The two latent heats on the other hand are not only quite different but huge in size compared to the specific heats. For example it takes more than 5 times as much energy to vaporize a gram of water as it does to raise its temperature from 0° to the boiling point. Thus phase changes are associated with considerable energy changes. If the phase change is up a phase then considerable energy must be supplied, if it is down a phase that same amount of energy will be released.

Melting (Fusion)	$L_m = 79.7 \frac{cal.}{gm}$
Vaporization (Condensation)	$L_v = 569 \frac{cal.}{gm}$

Why these latent heats are so large and why they are so different cannot be simply explained. In general however, it has to do with the way in which molecules bind together to form solids and liquids and the size of the forces which hold one molecule to its neighbors.

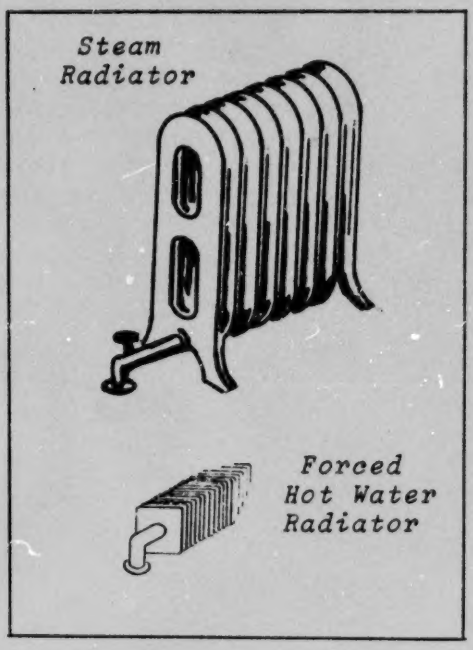
. . . . . AND THEIR SIGNIFICANCE:

Now the question remains "So what?". Other than helping us to better understand the nature and behavior of substances do these values have any practical significance. The question is an overwhelming yes, and in many rather profound ways. At the start of this module we saw that water was about the only substance that undergoes phase changes at temperatures near those that occur naturally on earth. Now we see that there is considerable energy involved in phase changes. Putting these two together and adding the plentiful abundance of water we find that there are many ways in which technology exploits these phenomena. Below are examples of only a few.

HEATING BUILDINGS: One of the basic concerns of man in northern climates is how to keep his buildings warm. One aspect of this is how to move energy around from room to room simply and cheaply.

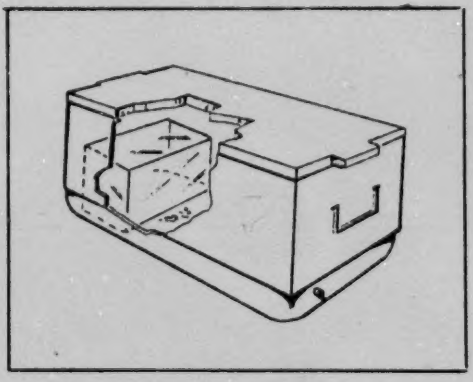
Steam is an obvious solution. It can be easily piped around, it carries large amounts of energy in its large latent heat of condensation, and the energy can be simply extracted by letting it condense to water in a radiator.

Since steam condenses at 100°C however, steam heating systems are quite hot and can be considered hazardous in the home. Thus forced hot water is being used more and more for domestic heating. While the specific heat of water is much less than the latent heat of steam, water is considerably more dense. That is when water is a liquid, many more grams can be circulated per second than when it is a vapor, and the total heat transported can be made the same but at a lower, safer radiator temperature.



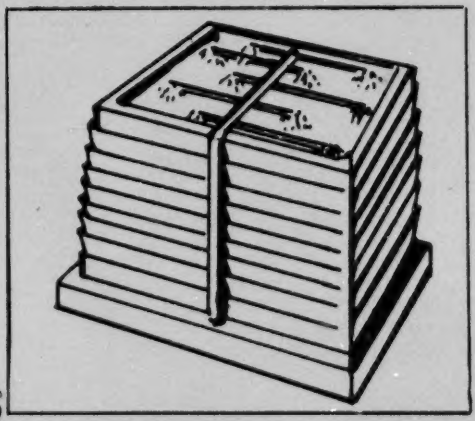
REFRIGERATION: In warm weather cooling is important. Again the latent and specific heats are exploited. The large latent heat absorbed by an ice cube melting compared to the small specific heat of a liquid in cooling means that it is very efficient at cooling liquids.

The old fashioned ice box as well as the modern picnic coolers depend on the large latent heat of melting ice. This latent heat is large compared to the specific heat of the food it is used to cool and thus provides temporary, constant temperature environment for storing food.



AIR CONDITIONING: When water evaporates into the air, this is a phase change much like boiling. (You will study this in more detail in Part 3). As in boiling the latent heat of vaporization must be absorbed for each gram of water that evaporates. The air conditioning system of large buildings utilize this principle in huge cooling towers inwhich water is made to flow over large surface areas exposed to the air.

As the water evaporates it carries the heat of the building in its large latent heat of vaporization.



THE SPECIFIC AND LATENT HEATS OF OTHER SUBSTANCES:

The quantities of specific and latent heat have been measured for each of the phases and phase changes for most other substances. However, since we generally encounter a substance in but one, or at most two, of its phases it is only the data for these that is of interest to us. For steel, for example, it is the specific heat of the solid that is important and no one cares much about its latent heat of vaporization or the specific heat of the vapor. Specific heats of a number of substances in their common phase are shown in the tables below. Since specific heat can vary with temperature we have given only the value for room temperature (unless otherwise indicated). For other temperatures consult the Handbook of Chemistry and Physics. Note also that no units are given for the numbers. That is because the specific heat of water is by definition the same for all the systems of units, namely 1.00 (1.00 BTU/lb°C, 1.00 cal/gm°C, 1.00 kcal/kg°C). Thus other specific heats are simply a certain fraction of the value for water and any of the units can be used with this fraction.

SPECIFIC HEATS OF COMMON SUBSTANCES

Phase	Substance	Specific Heat
	Ice	.50@0°C
	Aluminum	.21
S	Brass (60% Cu, 40% Zn)	.092
O	Copper	.092
	Earth (Dry Clay)	.22
L	Glass (approx.)	.11-.20
	Gold	.031
I	Iron (Steel)	.11
	Lead	.031
D	Silver	.056
	Stone (Granite)	.19
S	Wood (average)	.42
	Zinc	.093
	Water	1.00
L	Acetone	.53
I	Alcohol (Ethyl)	.58
Q	Ammonia	1.125
U	Carbon Tetrachloride	.20
I	Ethylene Glycol	.57
D	Mercury	.033
S	Olive Oil	.47
	Sulphuric Acid	.34
	(at Constant Pressure)	
	Hydrogen	3.39
V	Helium	1.3
A	Steam (@100°C)	.48
P	Nitrogen	.25
O	Air	.24
R	Oxygen	.22
S	Argon	.125
	Carbon Dioxide	.20

Also given are the latent heats for a few substances that exhibit phase changes near room temperature. Compare the values given with those you have measured for water.

LATENT HEATS OF COMMON SUBSTANCES

Phase Change	Substance	Temp. Of Change (°C)	Latent Heat Cal/gm)
	Alcohol (ethyl)	-114	24.9
M	Ammonia	- 75	108.1
E	Mercury	- 39	2.82
L	Water	0	79.7
T	Tin	232	13.8
I	Lead	327	5.47
N	Aluminum	658	94
G	Ethylene Glycol	- 12	42
SUBLIMATION	Carbon Dioxide	20	35.1
V	Water	100	569
A Z	Alcohol (Ethyl)	78	204
P A	Ammonia	-33.4	327
O T	Mercury	357	70.6
R I	Acetone	56	125
I O	Ethylene Glycol	197	191
- N	Freon		

MEASURING ENERGY AND ENERGY RATES (POWER)

What is energy? It is a property of a physical system, like a definite quantity of matter, that depends *only* on the condition of the system and not on how it got into that condition. If you want to change the condition of a body, you have to transfer energy from the outside, into the body. When this change is done mechanically the transfer process is called work. For this reason some books state that "energy is the capacity to do work". But you can change the condition of a body without doing mechanical work. In this module you are changing the condition of water by having energy transferred into it as heat flow. The melting of ice, the heating of water and its vaporization are all changes that are brought about by heat flow into the water.

Energy then is something stored in a body or a system and it can be stored in many forms such as deformation energy, chemical energy, electrical energy, magnetic energy, thermal energy, energy of motion, etc. It can be transferred from one body to another by one or more of the following processes: mechanical work, heat flow, flow of radiation. Energy is also *conserved;* it cannot be created or destroyed but only transferred from one system to another. Depending on what happens to the transferred energy, we say that energy is transformed in the process.

Energy is so important that we must find ways of measuring it, so that we can answer the question "how much?" There's no way of knowing how much energy resides in a body, but we can measure how much of it is transferred (or transformed) from one body to another. The measures and units we use depend, because of history, on the particular transfer process, but they can all be converted into a common set of units.

In summary, we measure changes of energy by the amount of work done, the quantity of heat transferred, or the radiant energy absorbed.

Power is *the rate at which energy is being* transferred, that is the amount of energy used per unit time. Very often power is defined as the rate of doing work. Since doing work and expending energy are essentially the same thing, the units which describe them are identical. When one says that something is powerful that generally means that it can do a lot of work. But it also means that it is using energy at a furious rate.

The word work has a slightly broader and more specific meaning in physics than that which we use in casual conversation. When we normally think of work we think of it in the context of some human or mechanical system. For example machines are often said to be able to do the work of X number of men. In physics, however, *work means the energy transferred when a force moves an object through a distance.*

When energy is tranferred from one system to another by mechanical forces (gravitational forces, pressure forces, or contact forces in general), the rate of energy transport is called mechanical power. If the forces are electrical, we talk of electrical power;

if chemical reactions result in energy transfer, we often talk of chemical forces and the work they do. Heat current (the quantity of heat transferred per unit time) is clearly a measure of power and might well be called thermal power.

In every case the kind of power we talk about refers to the means by which energy is taken from some source and delivered to something else. The loss of energy in the source or the gain of energy of the receiver is labeled mechanical, electrical, chemical, thermal, etc. depending on the changes that go on in these systems.

POWER AND TOTAL ENERGY. . . .

In most practical situations it is the rate of doing work (or expending energy) that is of interest. Thus most systems will specify their power rating, the amount of energy it uses per unit time. This is particularly true of electrical devices; electric motors, light bulbs, your immersion heater and so on, for which the power is listed in terms of watts. *One watt is a unit of electrical power equal to a rate of working of one joule per second.* If this rate is constant it is then a simple matter to obtain the total energy used over a given period of time;

$$\text{Total Energy} = \text{Power } \left(\frac{\text{energy}}{\text{time}}\right) \times \text{Time.}$$

In your electric bill for example you are charged so much per kilowatt hour (around 5¢) which is a measure of how much total energy you used over a period of time. Of course the units are a little strange since one kilowatt is 1000 joules/sec and you are multiplying this by the number of *hours*. But if this number is multiplied by 3600 sec/hour you would have the energy in joules (around .0014¢ per joule).

In your experiment you have determined the electrical power fed into the immersion heater. Since you will want to know how much total energy it took to produce various results (to melt all the ice; to raise the water from 0°C to 100°C; to boil away X number of grams of water) you can simply multiply the time it took times this input power.

CHOOSING THE RIGHT UNITS:

In this experiment we are concerned with energy that was produced electrically in the heater but transformed into thermal energy of the H_2O. Our calculation of electrical energy gave us energy in *joules*, a unit of energy used in the MKS system of electrical units. A joule is the amount of energy transferred per second when one ampere of current flows across a potential difference of one volt.

When one discusses the changes of substances caused by heat flow, however, these energy units are not generally used. Typical thermal units are the BTU and the calorie. And the definition of these two units are in part based on the thermal properties of water.

The calorie is defined as the amount of energy required to raise the temperature of one gram of water one degree Celsius.

The BTU (British Thermal Unit) is the amount of energy required to raise the temperature of one pound of water one degree Fahrenheit.

In analyzing your data you will compare your experimental results to these definitions.

ABOUT THE TABLE OF UNITS:

The relation between these and some of the other common energy units is given on the opposite page. Take a moment to familiarize yourself with the table and how to use it. Note also some of the following properties of the table and its units:

* The two units used most often in mechanical problems are the joule (newton-meter) and the foot-lb. These units are a force (newton, lb force) times a distance (meter, foot). Mechanical work is performed when a force moves an object through some distance. As indicated in the first paragraph doing work is a way of extracting the energy from one system and adding it to another.

  The joule is the amount of work performed (or energy dissipated) when a force of one newton moves a body through one meter in the direction of the force.

  The foot-lb is the amount of work performed when a force of one pound moves a body through one foot in the direction of the force.

* The joule, the calorie and the foot-lb are all about the same size measure of energy but in different unit systems. Obviously this amount is a convenient quantity in which to measure energy.

* The calorie that is referred to with regard to the energy content of food is in fact the kilo-calorie. The average American consumes about 3000 food calories per day. This is 3000 kilo-calories or 3,000,000 cgs calories per day. This is equivalent to the amount of heat required to raise 30,000 grams of water from 0°C to 100°C.

* The ev (electron volt) is the smallest of the energy units and is used most often in atomic and nuclear measurements.

  The electron volt is the amount of energy an electron gains when it moves across an electric potential difference of one volt.

A similar table of the Power Units and Conversion Factors is found on page 29.

ENERGY UNITS AND CONVERSION FACTORS

A \ B	Joule	erg	BTU	Calorie	Kilo Calorie	Foot-lb.	ev			Primary Areas Used	Unit System
Joule	1	10^{7}	9.48×10^{-4}	.239	2.39×10^{-4}	.738	6.25×10^{20}			Mechanical Electrical	MKS
erg (dyne-cm.)	1×10^{-7}	1	9.48×10^{-11}	2.39×10^{-8}	2.39×10^{-11}	7.38×10^{-8}	6.25×10^{13}			Mechanical	cgs
BTU	1055	1.05×10^{10}	1	252	.252	778	6.58×10^{23}			Thermal Chemical	English
Calorie	4.19	4.19×10^{7}	3.97×10^{-3}	1	10^{-3}	3.09	2.62×10^{21}			Thermal Chemical	cgs
Kilocalorie	4186	4.19×10^{10}	3.97	10^{3}	1	3087	2.62×10^{24}			Chemical Nutrition Biological	MKS
Foot-lb.	1.36	1.36×10^{7}	1.29×10^{-3}	.324	3.24×10^{-4}	1	8.47×10^{20}			Mechanical	English
ev	1.60×10^{-19}	1.60×10^{-12}	1.52×10^{-22}	3.82×10^{-20}	3.82×10^{-23}	1.18×10^{-19}	1			Atomic Nuclear Solid State	MKS

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HOW TO USE THE TABLE:

If you have a unit in column A and you want a unit in line B multiply by the factor shown.

FOR EXAMPLE: *If you have <u>ergs</u> and you want <u>BTU's</u> multiply by* $9.48 \times 10^{-11} \left(\frac{\text{BTU}}{\text{erg}}\right)$

$$\text{ergs} \times 9.48 \times 10^{-11} \left(\frac{\text{BTU}}{\text{erg}}\right) = \text{BTU}$$

POWER UNITS AND CONVERSION FACTORS

A \ B	Watt	Kilowatt	erg/sec	BTU/hour	Kilocal-orie/hour	Foot Pound /sec	Horsepower			Primary Areas Used	Unit System
Watt (Joules/sec)	1	10^{-3}	10^{7}	3.41	.858	.738	1.34×10^{-3}			Electrical	MKS
Kilowatt	1000	1	10^{10}	3410	858	738	1.34			Electrical	MKS
erg/sec	10^{-7}	10^{-10}	1	3.41×10^{-7}	8.58×10^{-8}	7.37×10^{-8}	1.3×10^{-10}			Mechanical	cgs
BTU/ hour	.293	2.93×10^{-4}	2.93×10^{6}	1	.252	.216	$3.93,10^{-4}$			Thermal Chemical	cgs
Kilocalorie/ hour	1.16	1.16×10^{-3}	1.16×10^{7}	.397	1	.858	1.56×10^{-3}			Thermal Chemical	cgs
Foot Pound/ second	1.36	1.35×10^{-3}	1.36×10^{7}	4.63	1.16	1	1.82×10^{-3}			Mechanical	English
Horsepower	745	.745	7.45×10^{5}	2544	642	550	1			Mechanical	English

29

HOW TO USE THE TABLE:

If you have a unit in column A and you want a unit in line B multiply by the factor shown.

FOR EXAMPLE: *If you have* Watts *and you want* Horsepower *multiply by* $1.34 \times 10^{-3} \frac{HP}{Watt}$

$$Watt \times 1.34 \times 10^{-3} \frac{HP}{Watt} = HP$$

Review

SUMMARY

The pressure cooker was chosen as the object of study because it is a useful and readily available device that illustrates some important principles of physics. In particular, one can learn about such features of matter as phases and phase changes.

Many (but not all) substances occur in three *phases* - *solid*, *liquid*, and *gas* (or vapor).

Phase changes involve energy changes: *melting* and *vaporization* require the addition of energy; *freezing* and *condensation* release energy.

Water is important and interesting for a number of reasons. Water is essential to living organisms. It is the only common liquid, and the only substance found naturally in all three phases.

Boiling and freezing temperature of water (and other substances) depend on pressure. The usual values apply only at one atmosphere pressure.

The first experiment of the module was an investigation of the energy required to heat substances and cause phase changes. Water was chosen as the substance because of the convenient temperatures of its phase changes.

A *thermistor* was used to measure the temperature of the water because its small mass does not disturb the temperature of the substance being measured. Thermistors respond to a temperature change by a change in resistance. Its negative coefficient of resistance means that as the temperature goes up its resistance goes down.

Calibrating the thermistor means finding the temperatures that correspond to various resistance values. For a thermistor, log R is proportional to 1/T, where T is *Kelvin* temperature. Such a relation is conveniently plotted as a straight line on *semilog* graph paper.

R vs T can also be plotted on semilog paper, to give a more convenient *calibration curve*. Many points are needed to determine the curve accurately.

Energy was supplied to the water electrically by an immersion heater.

Energy is a physical property that can be stored in a body. It can be stored in several forms: *deformation* energy, *chemical* energy, *electrical* energy, *magnetic* energy, *thermal* energy, *kinetic* energy,etc.

Energy is transferred from one body to another by *work*, *heat flow* or *radiation*. In this experiment it was done by heat flow from the heater to the water.

Power is the rate of transferring energy from one body to another. When energy is being transferred at the rate of one *joule per second*, the power is one *watt*. Electrical power input to the immer-

sion heater is found by measuring applied *voltage* and heater *resistance* and using the relation:

$$P \text{ (watts)} = \frac{V^2}{R}$$

When energy is being supplied at a constant rate (constant power), the *total* energy supplied is:

$$\Delta Q \text{ (joules)} = \frac{\Delta Q}{\Delta t} \text{ (watts)} \cdot \Delta t \text{ (secs)}.$$

There are two important thermal properties of matter: *specific heat*, and *latent heat*.

Specific heat is the energy required to raise the temperature of a *unit mass by one degree*. The specific heat of water defines the *calorie* and the *BTU*.

Latent heat is the energy required to produce a phase change in a unit mass.

Latent heats are very large compared to specific heats. This is because phase changes involve major rearrangements of the atoms and molecules of a substance.

The large latent heats for water have many applications in heating and refrigeration systems.

QUESTIONS:

1. Pick out which of the following units are *not* units of energy:
   a) BTU
   b) calorie
   c) watt
   d) erg
   e) joule
   f) horsepower
   g) foot-pound

2. The Chemical Rubber Company (CRC) book of tables is a useful and convenient source of many kinds of information.
   Look up the melting point and boiling point of the elements, magnesium and cesium.
   Plot these values in the blank spaces on the graph of page 8.

3. When heating the water in the experiment of this part, the temperature rose steadily to 100°C. The temperature then stopped rising.
   Explain briefly why this happened.

4. How is *power* different from *energy*?

5. What is the difference between specific heat and latent heat?

PROBLEMS:

For the following problems refer to the data in the tables on page 24.

1. Calculate the heat required (in calories) to raise the temperature of 100 gm of brass from room temperature (20°C) to its melting point (900°C).

2. Ten grams of ammonia gas condense to form liquid ammonia. How much heat is released?

3. A certain electric motor is rated at one horsepower. How many ft-lb of work can it do in two minutes?

4. A certain lamp uses electric energy at the rate of 250 watts. How long will it take the lamp to use a kilowatt-hour?

5. How many joules of energy are required to raise the temperature of a cup (250 gm) of water from 20°C to 100°C and then boil it away?

THE GAS PHASE

INTRODUCTION

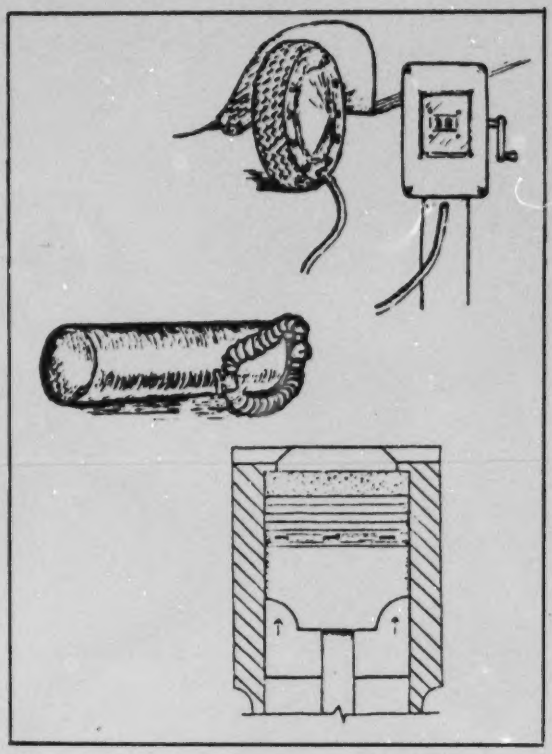
Here in Part 2 of this module you will use a household pressure cooker to help you learn about some important features of the behavior of gases. In Part 3 you will examine the pressure cooker more closely and see why it is useful in the home, the medical laboratory, and in industry. Right now you make use of the cooker because it is a convenient device that can be made air tight and can stand fairly high temperatures and pressures.

Many thousands of natural effects and familiar devices can be understood only in terms of the behavior or properties of the gaseous form of matter. Thunderstorms, hurricanes, and our daily weather depend upon the properties of the gases air and water vapor. That is, what we call weather is really a constantly changing set of values for air temperature, air pressure, air volume, and quantity of water vapor.

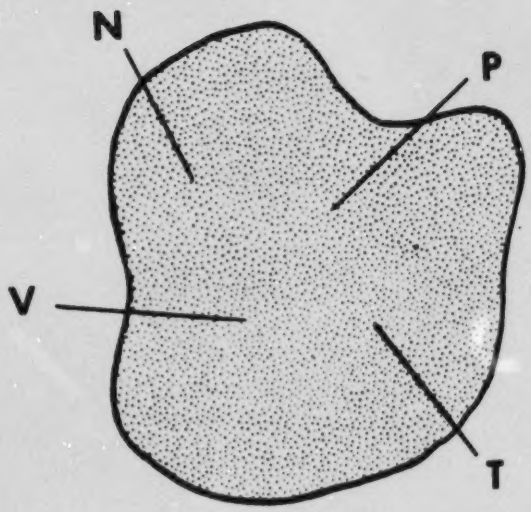
THE FAMILIAR BEHAVIOR OF GASES. . .

The usefulness of an automobile tire depends in part upon the fact that cramming more air into it raises the pressure. The standard scuba air tank is said to have a capacity of 71.2 ft^3 of air, but actually it is a cylinder only 6 inches in diameter and 24 inches long. The volume of such a cylinder is less than half a cubic foot. How can it hold more than 71 ft^3 of air? You know the answer: 71 ft^3 of air at ordinary atmospheric pressure can be crammed into the small tank, but the pressure becomes enormously high, about 2200 pounds per square inch, or 150 times normal atmospheric pressure.

In the cylinder of a diesel engine the fuel-air mixture is compressed to a small fraction of its original volume by the piston; this reduction in volume raises the pressure and also raises the temperature so high that the mixture burns spontaneously without a spark plug.



. . .IS DESCRIBED BY THE GENERAL GAS LAW:



In all these situations, and many others, we are concerned with what is called the "state" of the gas. When you ask the question "what is the state of a gas?", you are really asking a number of questions: "how much gas do I have, (its mass or number of molecules?"), "what is the pressure?", "what is the temperature?", "what is the volume?". There can be other quantities to consider, but *number of molecules*, *pressure*, *volume*, and *temperature* are the key factors that describe the state of a sample of gas.

Our aim in this part then is to investigate the relation among these quantities; to see how N, P, V and T are related. This relation is summed up in an equation called the General Gas Law or Ideal Gas Law. In the experiment and the analysis that follows you will study quantitatively the behavior of a gas and build a statement of the General Gas Law based on your measured data.

THE GENERAL GAS LAW: OR

WHY CALL THE LAW GENERAL?

In the simple description of a gas it is only the *number* of molecules, the volume, the temperature, and the pressure that we need to be concerned about. These are called the *parameters* which describe the gas. The General Gas Law is an equation which expresses the relation among these four parameters. The law is called "General" because it does not depend on the type of gas you are talking about. Regardless of what the gas is or, even whether it is a mixture of several gases, the law can be applied equally well. For example, air is composed of 78% nitrogen, 20% oxygen, and 2% other gases, as well as varying amounts of water vapor. If each one of these substances behaved differently with respect to a change in say the temperature, it would be an extremely difficult task to predict how the air as a whole would behave. However, since they all respond the same, the General Gas Law can be simply applied to the mixture, whatever it is.

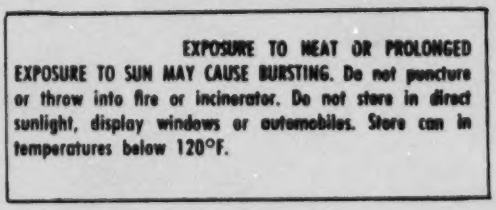
. . .AND WHEN CAN IT BE USED:

Now that we have explained the virtues of the General Gas Law we should add a note of perspective. Other names for the Law are the "Perfect" Gas Law and the "Ideal" Gas Law. These names are used because the Law is not accurate over all the values of P, V and T that a gas can have. For example, just before a gas starts to condense to form a liquid it gets quite dense, the molecules are quite close together, and this gas law no longer works. Only for an idealized or perfect gas would the General Gas Law hold under all possible conditions.

But provided you are not too near the condensation temperature or at extremely high pressures these differences are not important. The General Gas Law can therefore be used for most situations you will be likely to encounter. That is, it can be used to predict the behavior of a gas, any single gas or mixture of gases, when its number of molecules, volume, pressure, or temperature are changed.

FOR EXAMPLE:

Temperature and closed containers: Any time a closed container is exposed to heat, the pressure inside increases. If the pressure exceeds the strength ratings of the container then the container may burst causing a hazard to anyone around. The warning opposite is a typical one found on closed cans, such as aerosol cans, to caution the consumer of this danger. The


EXPOSURE TO HEAT OR PROLONGED EXPOSURE TO SUN MAY CAUSE BURSTING. Do not puncture or throw into fire or incinerator. Do not store in direct sunlight, display windows or automobiles. Store can in temperatures below 120°F.

N, V, T, P AND THEIR RELATIONS:

General Gas Law can be used to calculate the expected pressure in such *closed gas-filled containers* due to a given temperature increase. (Note, it cannot be used for a *full* aerosol can since they are filled with a high pressure liquid. How to do this will be seen in Part 3).

Pneumatic systems: Pneumatic systems depend on the force of high-pressure or compressed air to operate their moving parts. To "compress" air means to decrease its volume. The amount of volume reduction can be simply calculated from the General Gas Law.

FIXED vs VARIABLE PARAMETERS:

In this section we want to experimentally determine the relation among N, V, T, and P that describes the general behavior of gases. But this is not a simple single experiment. After all there are four parameters. If we change one, all the other three may change and we would have to carefully measure and separate out each one of these changes. Then we would have to change a second parameter and again measure changes in the other three and so on. This would be a difficult, if not impossible, experiment to carry out.

A much simpler way is to *fix* two of the parameters, that is be sure they cannot change, and then determine the relation between the remaining two; (change the third and see how the fourth varies). Then fix two different ones and determine the relation between the other two and so on until the relations between all of the pairs of parameters have been determined.

YOUR EXPERIMENT:

Clearly this would be a long set of experiments and far too much to do in a single laboratory period. However, they all have been done and the relations determined. What you will do is use the pressure cooker as a container with a *fixed volume* and it will be carefully sealed so that no air molecules can get in or out, that is there will be a *fixed number of molecules*. You will then heat up the pressure cooker, *changing the temperature* of the air inside (measuring it carefully with your thermistor) *and measure the variations in pressure* of the air.

Having performed this experiment and analyzed the results we will then discuss qualitatively other relations from experiences with which you are familiar. From this we can then give an expression for the General Gas Law.

A MEASUREMENT OF P vs T...

In this experiment we want to be sure that the mass of gas and its volume remain constant as we vary the temperature. The pressure cooker makes an ideal container for meeting these conditions. First it comes with a rubber gasket so that the chamber may be sealed and no gas can leak in or out during the experiment. One of your first tasks will be to check that seal to be sure that it is leak tight. Second, it has a fixed volume. Being made of about 3/16" thick cast aluminum, it won't expand significantly during the measurement. Thus we have confidence that N and V will indeed be fixed.

Next we must be concerned about measuring the temperature of the air inside, and its pressure.

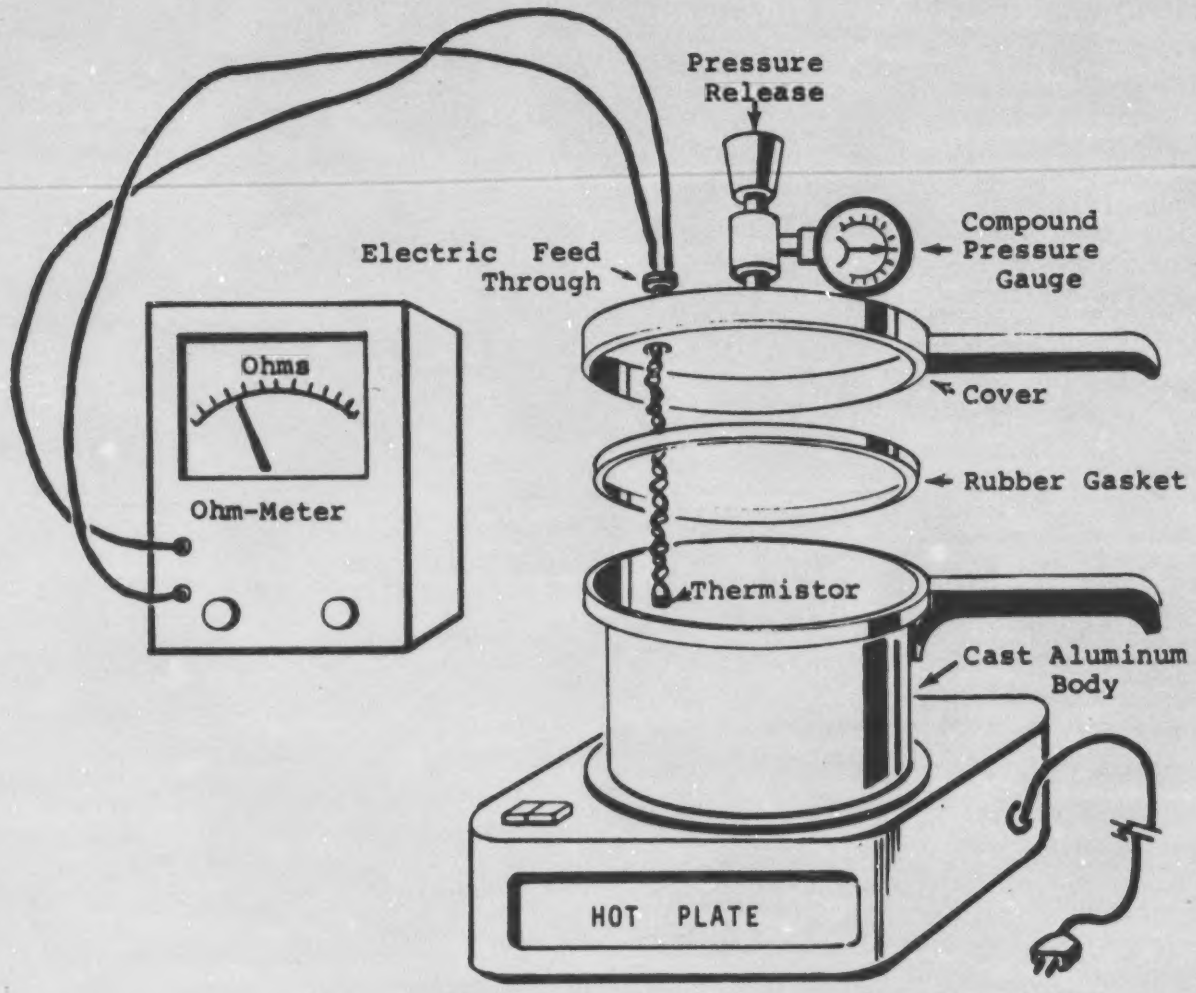
MEASURING THE TEMPERATURE:

To measure the temperature you will use the thermistor that you calibrated in part 1. The pressure cooker has been fitted with an insulated, pressure tight, electrical feed through so that you can attach the thermistor inside the cooker but measure its resistance from outside. In attaching the thermistor, however, be sure that it does not touch the container walls. The container may become hotter than the gas during heating and damage the thermistor or the thermistor may short out during the measurement and you will have to open up and start over.

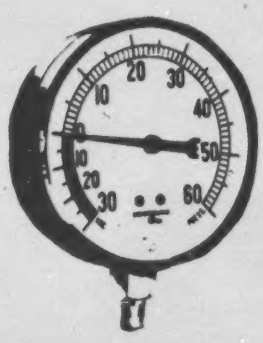
MEASURING THE PRESSURE:

The pressure gauge attached to the pressure cooker is a compound gauge, meaning that it can be used for measuring pressure above atmospheric and below atmospheric (or vacuum). Since atmospheric pressure is marked 0, you will be measuring *gauge pressure*, not absolute pressure. We will discuss this difference later, but essentially what it means is that your readings will all be relative to the atmospheric pressure of the day. Therefore, you must determine from a barometer the atmospheric pressure at the time of the readings.

To show you how confusing units can get, the divisions on the pressure gauge scale are different. Those above 0 are in terms of pounds per square inch (lbs/in^2) while those below are in terms of inches of mercury (in/Hg). This is a common convention in gauges so they don't bother to mark the units on the scale. Again we will discuss the units later but be sure you record which units you are reading for later conversions.



Compound Pressure Gauge

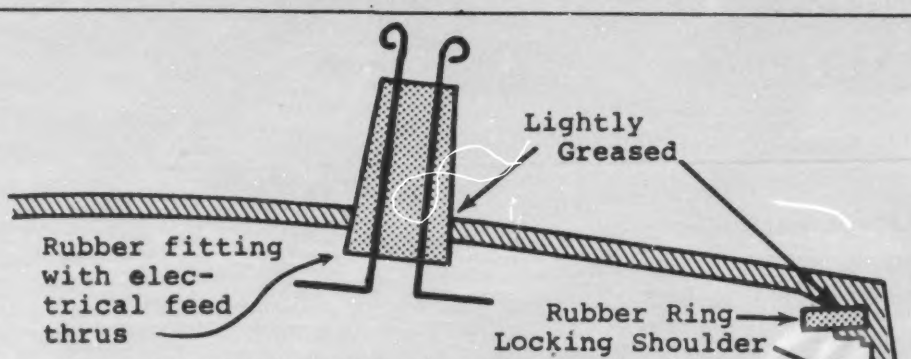


<u>The Experimental Set Up</u>: *This is a schematic diagram of a pressure cooker modified for measuring the P-T relation of a gas. The cooker is sealed by the rubber gasket. The hot plate is used to raise the temperature of the gas which is measured using the thermistor. The compound pressure gauge simultaneously reads the gas pressure. A pressure release is provided in case of accidental overpressure in the chamber.*

SOME SIMPLE EXPERIMENTS:

ASSEMBLE THE APPARATUS

Assemble the pressure cooker to make an air tight container. It must be put together properly to get a good pressure seal. Remove the gasket and the fitting for the electrical leads. Make sure they are uniformly and *lightly* greased and then replace them securely. Check that the threads on the pressure gauge fitting are tight. When you have completed these preliminary checks, close the pressure cooker carefully and be sure the cover is locked in place. We will then try some simple experiments to get a feeling for the physical parameter of pressure.



CROSS SECTIONAL VIEW OF THE COVER. The rubber fitting and ring gasket should be removed and coated with a uniform but fairly thin layer of vacuum grease.

TEST YOUR LUNG POWER

First we will use the gauge on the pressure cooker to test your lung power. Place a rubber hose on the pressure release hole. Blow in and see how much pressure you can produce. Caution: *This is only a rough test so do not try to burst your lungs.*

Suck out and see what vacuum you can produce.

Are your lungs more effective in producing vacuum or pressure?

Remember that the scale is calibrated in two kinds of units:pressure is in lb/in^2, and vacuum is in inches of mercury. Roughly speaking, one lb/in^2 is the same as two inches of mercury. Record your values.

Lung Pressure_____Lung Vacuum_____

* TEST OTHER THINGS. . .

Using the pressure release hole as an access to the pressure gauge test the vacuum produced by some other common vacuum devices. For example try an aspirator, a vacuum cleaner, the manifold vacuum of your automobile carburetor, etc. Do you know the principles behind which these devices work? Caution: *Do not test for pressures over atmospheric unless you are confident they are less than 15 lbs/in^2. As you will see later, the forces on the lid become tremendous and could be potentially dangerous!*

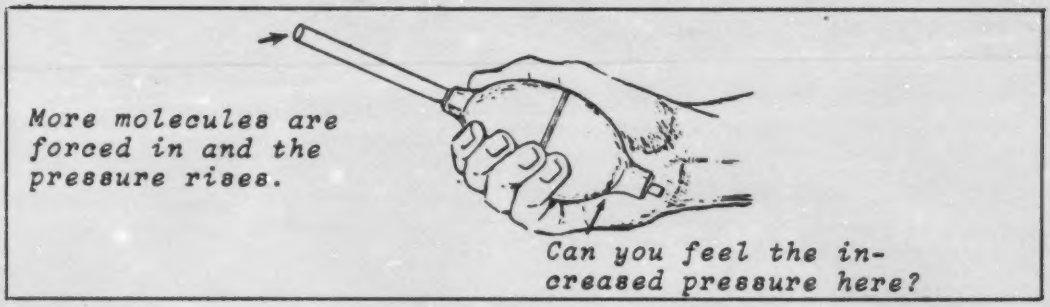
USE THE SQUEEZE BULB

Connect the squeeze bulb to the pressure cooker. What pressure can you build up? Pay particular attention to the way the bulb feels; do you believe that pressure and force are closely related?

Every time you completely squeeze the bulb you are forcing in the same number of molecules ΔN. How much does the pressure change (ΔP) for each squeeze? Since V and T are fixed what does this say about the relation between N and P?

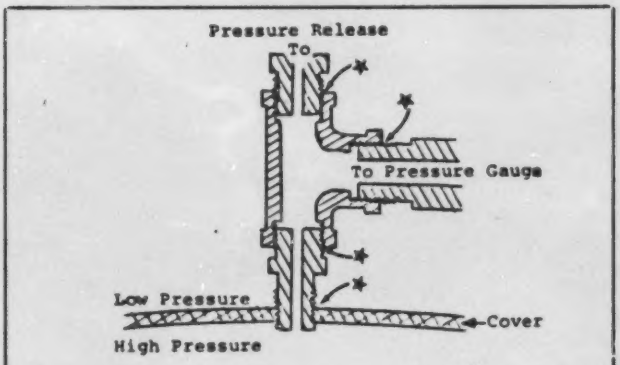
Turn the bulb end-for-end, and connect it "backwards". Can you produce a vacuum? How much? Why is there a definite limit to the vacuum you can draw this way?

Pressure_____ Vacuum_____



TEST FOR LEAKS

The leak test is a modern version of an old procedure used for checking automobile inner tubes before the days of tubeless tires. Dunking an inflated inner tube in a tub of water and watching for bubbles was a very sensitive test for leaks. You will use nearly the same trick here. Using the rubber bulb, pump the chamber up to a pressure of about 5-10 lbs/in^2. Apply soapy water (or saliva) to all possible places where the container could leak: the threads and fittings for the gauge, etc; around the gasket seal between the cover and the container; around the ohmmeter leads and their rubber holders. A leak is indicated by the presence of bubbles.



Apply soapy water to all threaded joints ★. If bubbles appear, there is a leak and the joint must be tightened.

If you see any growing bubbles, try to eliminate the leak. If it is at a threaded joint, tighten the joint. If the leak is around the electrical feed thru, apply more vacuum grease. If there are leaks on the main cover seal, be sure that the gasket is properly seated and coated with a uniform, but not too thick layer of vacuum grease. Caution: *Be sure to release the pressure before removing the cover!*

THE STEPS

AN OVERVIEW OF THE PROCEDURE

The purpose of this experiment is to obtain the relation between P and T for a gas. You will do this at chamber pressures below atmospheric. If the inside pressure is less than that outside, then the force of the atmosphere will be pushing down on the cover. It will be pressing on the rubber gasket and helping to make a good pressure-tight seal. If the experiment was performed at pressure in the chamber greater than that outside, the gasket would probably still seal (as your leak test showed) but in general it is best to have atmospheric pressure working with you rather than against you.

FIRST PREPARE A DATA TABLE. . . . .

As usual you will need a table to record your data. Your table should look similiar to that shown at the right. In the column for Notes be sure to record any actions you took in the procedure (such as adding a fan). Have your thermistor calibration graph available for quick reference.

Notes	R (ohms)	T (°C)	P (in Hg)		

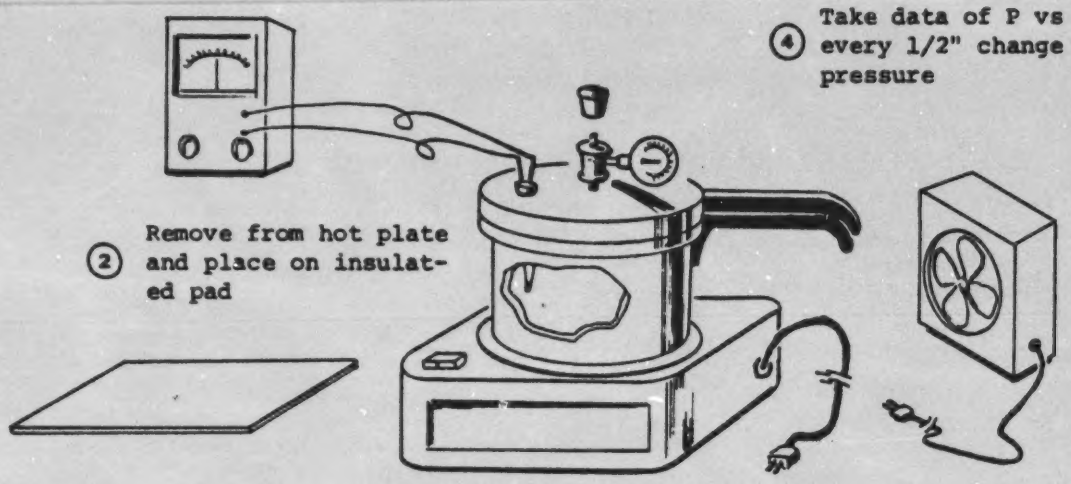
. . . . .THEN PERFORM THE EXPERIMENT:

1. Place the pressure cooker on the hot plate and start heating. Leave the pressure release off so that the hot gas can escape. Continue heating until the thermistor registers a temperature of about 160°C. Caution: *Don't exceed this temperature or the solder in the electrical connection will melt.*

2. When the temperature reaches 160°C remove the cooker from the hot plate and set it on a heat-resistant surface. Caution: *160°C is very hot.* Initially the bottom is the hottest surface and you must wait until the cooker body and the gas inside all come into thermal equilibrium. How do you know when equilibrium is reached? You can't be sure exactly but a good indication is when the thermistor resistance stops changing rapidly. This should take a minute or two.

3. When the thermistor stops changing, seal off the chamber by replacing the rubber stopper. Take an initial reading of pressure and thermistor resistance. The pressure, of course, should be 0.

4. Continue taking readings of resistance and pressure as the chamber cools. Note: *Pressure gauges that you are using have a tendency to stick. Before each pressure reading tap the gauge lightly a couple of times.* Take readings at each half inch change in pressure. This will allow you to write down the pressure readings immediately and then simply record the thermistor resistance as that pressure is reached.


③ Replace pressure release to seal off container
④ Take data of P vs R every 1/2" change in pressure
② Remove from hot plate and place on insulated pad
① Heat pressure cooker to 160°C with pressure release off
⑤ Blow fan on chamber to speed the cooling

By working at a chamber pressure below atmospheric, the force of the atmosphere will press down on the cover and help to insure a good seal of the gasket.

5. The cooling process will go slower and slower as the temperature falls. Do you know why? You can speed it up by using a fan to promote heat losses by forced convection. Note: *Be sure to stop the fan somewhat before the gauge pressure reaches the pressure you want to read so that the system has time to come into thermal equilibrium.*

* 6. Optional:

Once the cooker has reached room temperature you may want to take additional data below room temperature. This can be done by placing a wet cloth on the cooker with a fan on it for evaporational cooling, placing the cooker in cold (or ice) water, or outside on a cold day, etc.

SOME FINAL DATA:

For later calculations you will need data on the ambient pressure and temperature, that is the pressure and temperature of the air surrounding the experiment. Find the values and record them.

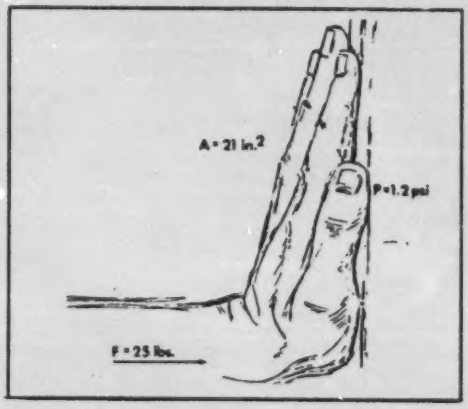
Temperature_____ Pressure_____

PRESSURE AND FORCE

Before you begin to analyze your data we want to discuss the parameter of pressure so that you will have a good feel for what is going on in the pressure cooker. The most important relation is that between pressure and force. That distinction is critical in the pressure cooker and in many other devices in which fluid pressures are involved such as hydraulics and pneumatics. Using the force-pressure relations properly has given great power to man and when not properly understood has led to disasterous accidents. We will also try to classify the many different units that are used to measure pressure.

HOW ARE FORCE AND PRESSURE RELATED. . . . .

As an example of how pressure and force are related let us consider your hand. Suppose you can exert a push on the wall of about 25 lbs. (You might check this value by putting a bathroom scale against the wall and pushing on it). This is a force of 25 lbs. Force as you will recall is a vector quantity meaning that force has a direction as well as a magnitude. The direction is perpendicular to the wall and its magnitude is 25 lbs. Now let us assume that the palm of your hand is about 3 in. wide by 7 in. long or an area of 3 in. x 7 in. = $21 in^2$, and that you spread the force out evenly over over your palm's surface. The 25 lb force is therefore spread over an area of $21 in^2$. The force exerted on each square inch of wall is only 25 lbs/$21 in^2$ = 1.2 pounds per square inch. This quantity is called the *pressure* on the wall, the number of pounds of force exerted on each square inch of surface. Or more generally, *the pressure is the force per unit area of surface.*

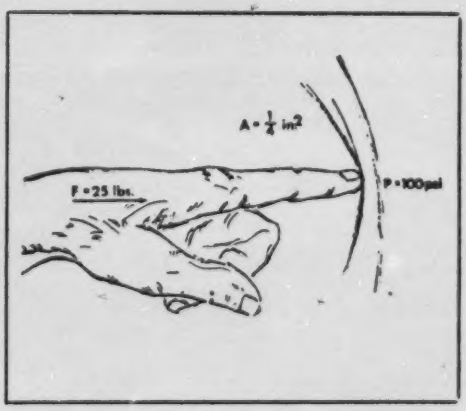


$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

$$P = \frac{F}{A}$$

$$P = \frac{25 \text{ lbs.}}{21 \text{ in}^2} = 1.2 \frac{\text{lbs}}{\text{in}^2} = 1.2 \text{psi}$$

Now suppose instead of putting your palm against the wall you put only one finger. The end of your finger is roughly 1/2 inch square or an area of 1/2 in. x 1/2 in. = 1/4 in^2. If your arm still exerts a force of 25 lbs., the pressure on the wall is:



$$P = \frac{25 \text{ lbs.}}{1/4 \text{ in}^2.} = 100 \text{psi}$$

Thus for the same force you can exert pressure anywhere between 1 and 100 psi.

PRESSURE IN FLUIDS - - Gases and Liquids:

Normally when you encounter pressure it will be in connection with a fluid, either a gas or a liquid, rather than in a solid body as in the example we gave on the opposite page. Solid objects resist forces which try to change their shape or volume. Fluids flow easily and hence change shape readily; liquids offer considerable resistance to compression, but it is relatively easy to change the volume of a gas. When a fluid is compressed you say that it is under pressure and this pressure exists everywhere in the fluid. For example, when you squeezed the bulb you exerted a force on it and the gas inside compressed, increasing its pressure, until it produced a force on your hand equal and opposite to the force you applied.

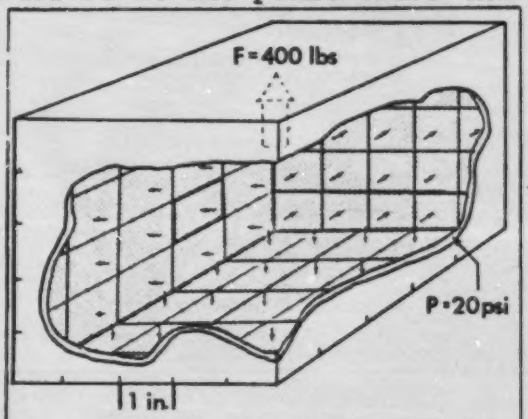
Pressure, then, gives us a measure of how hard a fluid pushes (or the force) on the surfaces that it contacts (as well as how hard one part of a fluid pushes on the part next to it). The forces which are pushing on the fluid, and the forces that the fluid exerts back, are called *pressure forces*. Since forces are vectors which have both magnitude and direction we must give a direction to these forces. The direction is determined by saying that *pressure forces always act at right angles to the surface* that the fluid touches, no matter what the direction of the surface. Thus fluid pressure is not a force but a condition that gives rise to a force. This force has a direction fixed by the surface on which it pushes and a magnitude that depends on the pressure at the point where the surface lies.

The measure of pressure is the magnitude of the pressure force that will be exerted on a square unit of surface. That is, when the pressure is uniform (the same everywhere in the fluid) each square unit of surface touched by the fluid has a perpendicular force on it whose magnitude is equal to the value of the pressure. If the pressure is 20 lbs/in^2, then every square inch of surface has a perpendicular force on it of 20 lbs. The total force on a surface will be the product of the total surface area times the pressure:

$$F = P \times A$$

where the force will always be at right angles to the area chosen.

But how do you determine the force on curved surfaces like those found in the pressure cooker? On the following pages we discuss the force on three of these, the top, the pressure gauge, the pressure regulator, and show how they may be determined.



A box filled with a fluid at a pressure of 20 psi has a force on each square inch of 20 lbs. This so called pressure force always acts at right angles to the surface regardless of the surface shape or orientation. The total force on a surface is the product of the pressure times the surface area. Thus the total force on each side of the cube shown above is:

$$F = 20 \frac{lbs.}{in^2} \times 25\ in^2$$

$$= 400\ lbs.$$

FORCES IN THE PRESSURE COOKER:

THE TOP: The surface of particular interest in the pressure cooker is the cover. We want to know the force on the cover so that we know how strong the cover latch must be. As seen in the diagram below, the cover is dome-shaped and the forces on it are not all acting vertically to push it off. Some of the forces have sideways components.

The total vertical force on the lid, however, can be simply calculated by asking what would be the force on the lid if it were flat. That is, we must choose a flat area that is at right angles to the force direction we are concerned about. A flat lid would have an area equal to the area of the top of the pressure cooker body, and the force on it would be equal to the product of that area times the pressure. For example, suppose the diameter of the top were 8.0 in. Then the area would be:

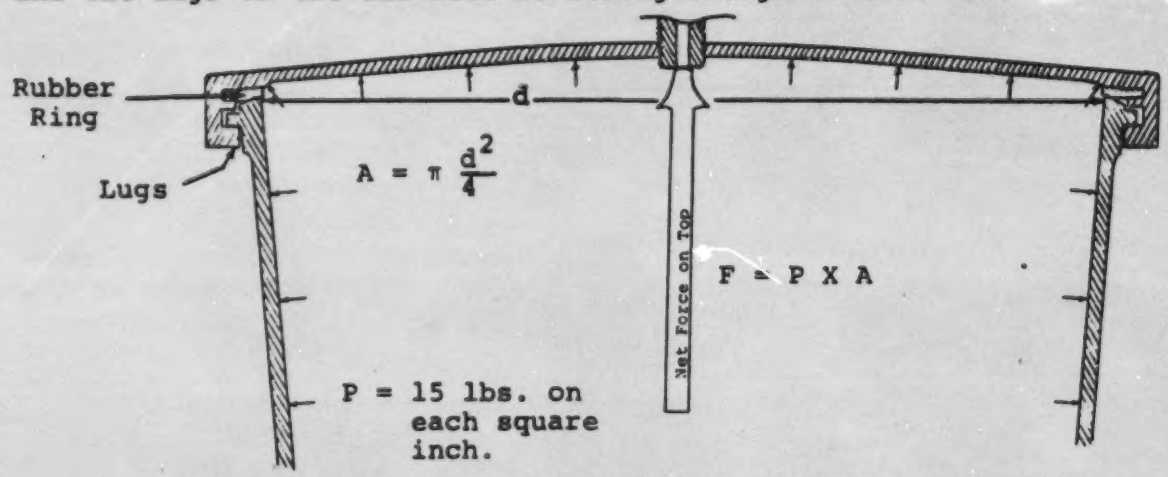
$$A = \pi \frac{d^2}{4} = 3.14 \times \frac{(8.0)^2}{4}$$

$$A = 50 \text{ in}^2$$

If the pressure in the cooker were 15 lbs/in^2, the total force on the lid would be:

$$F = P \times A = 15 \text{ lbs.} \times 50 = 760 \text{ lbs.}$$

Thus the force on the lid is more than 1/3 ton (2000 lbs./ton) and the lugs on the lid must be strong enough to hold this force.

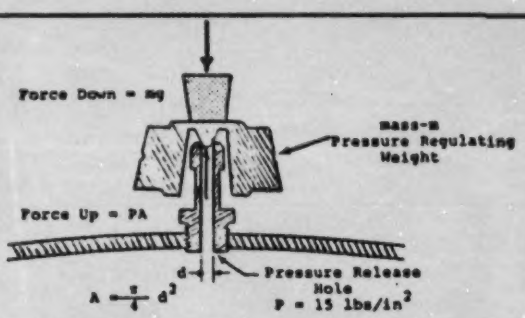


The force attempting to blow off the lid is simply the cross sectional area of the top times the pressure in the chamber. This force must be safely held by the lugs on the top and body of the pressure cooker.

THE PRESSURE REGULATOR: The pressure in a pressure cooker is regulated using the same principle. A small weight is fitted over the top. The area of the hole and size of the weight are chosen so that when the pressure inside reaches some definite value it will exert enough force to overcome the gravitational force and lift the weight off the hole and release the pressure.

Suppose we want the pressure to release at 15 lbs/in^2 and we choose the hole to be 1/8 in. in diameter (small enough not to require a large weight but big enough that it won't clog with food).

How heavy must the weight be?



The pressure-regulating weight fits snugly in the pressure release hole. It provides a good pressure tight seal until the upward force exerted by the pressure exceeds the gravitational force down.

The hole is:

$$A = \pi \frac{d^2}{4} = .012 \text{ in}^2$$

The force on the weight at 15 lbs/in^2 is:

$F = 15$ lbs. x .012

$F = .18$ lbs.

Thus the regulator must weigh 3 oz. (16 oz./lb.). You will see in the next section how effective this technique is in maintaining a constant pressure.

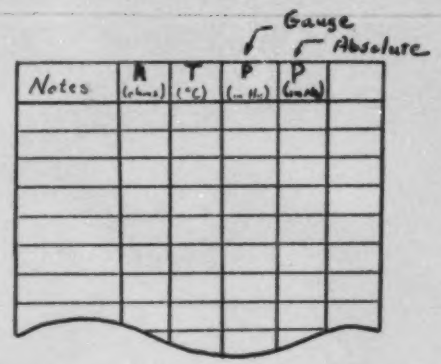
THE PRESSURE GAUGE: To read the pressures in this experiment you used a gauge called a *Bourdon gauge* The key to this gauge is the *Bourdon tube* which converts pressure changes directly into a mechanical motion that moves the needle. The tube is in the form of an arc and as the pressure increases it uncurls much like a child's toy that uncurls and shoots out as you blow. To *calculate* the force-pressure relation for this instrument is quite difficult so that in general the relation between force (or motion) and pressure is determined experimentally. Thus the scale markings are specially calibrated for each tube. One of the advantages of the Bourdon tube is that the relation is extremely linear (a plot of needle motions vs pressure is a straight line) and reproducable (it always shows the same value for the same pressure).

The Bourdon type pressure gauge works much like a child's toy. As the pressure in the tube increases, it uncurls. A mechanical system transforms this motion to the rotational motion of a needle which moves across a scale calibrated in pressure units.

ANALYZING YOUR DATA....

As usual, after collecting data from an experiment you must analyze the results and display them so that their meaning is clear. In this case we want to know the mathematical relation between pressure and temperature and a graph of these two will illustrate their relation nicely. Follow the procedure below to obtain your graph.

1. COMPLETE THE DATA TABLE.



The first step is to convert the data that you have into pressures and temperatures of convenient units. Using your thermistor calibration graph convert your resistance values into temperatures.

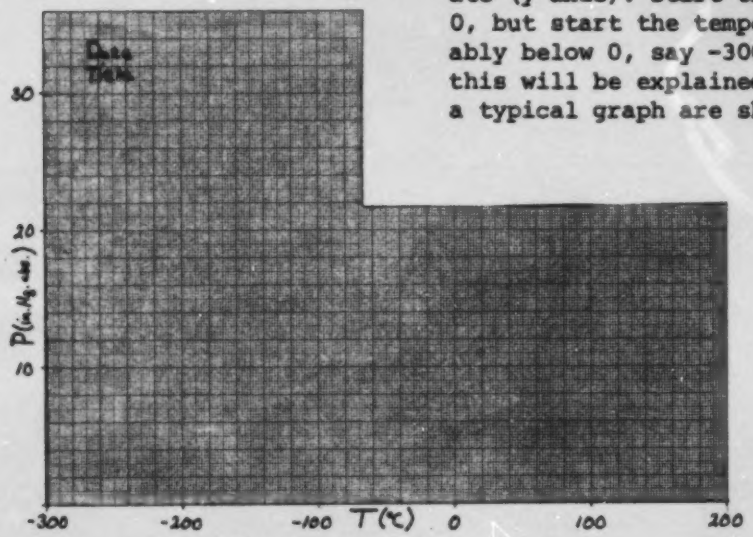
Next you must convert the gauge pressure readings of vacuum into absolute pressure values. If you are unfamiliar with the various units for pressure and the difference between absolute and gauge pressure, refer to the summary pages on the topic at the end of this section. There it is seen that:

P (absolute) = P (atmospheric) - P (gauge vacuum)

Using the value of atmospheric pressure that you determined from the barometer, convert your gauge pressure to absolute. Express your results in in.Hg.abs.

2. PREPARE THE GRAPH PAPER.

In setting up the graph of P vs T, let T be the abscissa (x-axis) and P the ordinate (y-axis). Start the pressure axis from 0, but start the temperature axis considerably below 0, say -300°C. The reason for this will be explained later. Portions of a typical graph are shown as an example.



3. PLOT THE DATA.

Now plot your P-T data points on the graph. You will find that it only fills a small portion of your graph.

4. DRAW THE GRAPH.

Your data points should fall pretty nearly on a straight line. Draw the best straight line you can through the points. In general this means that the number of points that are off the line above equals the number that are off below.

WHAT DOES THE GRAPH MEAN?

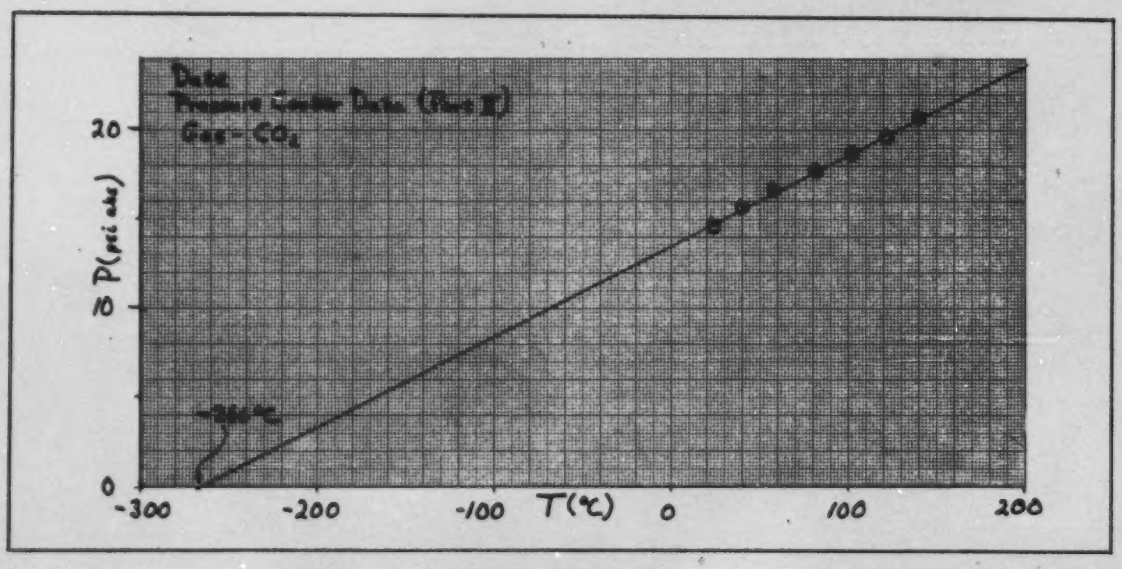
Since the graph of pressure against temperature is a straight line, the relation between absolute pressure and Celsius temperature is linear. This is not the same as saying that these two quantities are proportional. If they were proportional to one another, then the pressure would be zero at 0°C; this is obviously not true here. Also, when quantities are proportional, if you double one, for example, then the other one is doubled also. Look at how the pressure changes from say 50°C to 100°C; it certainly does not double. To answer the question raised above let us now ask the following question.

WHAT IS THE TEMPERATURE AT ZERO PRESSURE?

The lowest absolute pressure we can get is zero. What happens to the temperature then? We cannot get a really good answer without actually taking measurements to see what does occur at lower and lower pressures and temperatures. But we can get a useful general idea by doing what is called *extrapolation*. Here, this means starting with the straight line you have drawn on the graph and extending it backwards toward zero pressure. Below is a graph of sample data taken with the pressure cooker but using a different gas, carbon dioxide (CO_2), and over a different pressure region. On this graph the line when extended hit 0 pressure at about -266°C.

5. EXTRAPOLATE YOUR GRAPH TO P=0.

What temperature does P = 0 correspond to?_____.



What does this mean? If zero is the lowest possible pressure, then it may make sense to say that this temperature is the lowest possible temperature. But do we get this temperature only in this particular set of circumstances - only with this same amount of air, let us say? When the experiment is repeated many, many times with all kinds of gases, and different amounts of gas, the results are always the same. With careful measurements it turns out that when the line is extrapolated to zero pressure, this always occurs at *-273.16°C.*

ABSOLUTE ZERO. . . . .

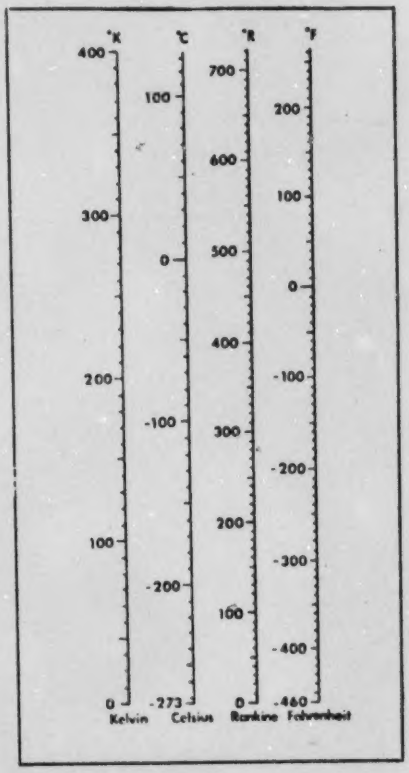
Since this number occurs over and over again, it makes sense to think that this really is a special temperature. Perhaps it actually is the lowest possible temperature. The evidence we have here is not positive proof. For one thing, extrapolation is always dangerous, since we are extending the results of measurements into a region where the results may not apply. Also, we know that all gases become liquids and most become solids long before reaching a temperature of -273°C. Therefore, it is hardly correct to speak of gas pressure becoming zero when the substances are in fact no longer gases.

In spite of these difficulties, other evidence does indicate that -273°C is indeed a special temperature and the lowest attainable temperature. Perhaps the most direct evidence for this is laboratory experiments of many different kinds in which the aim is to make the temperature as low as possible. All such experiments show that the lower limit is -273°C.

. . . . . AND ABSOLUTE SCALES OF TEMPERATURE:

This temperature is given the name of *absolute zero* since it is absolutely as low as temperature can get. Using this as the 0 of temperature and choosing the same size degree as °C, we define a new temperature scale called the *Kelvin* scale (°K). The Kelvin scale is thus an *absolute temperature scale.* It is much like the absolute pressure scale which had its zero at the lowest absolute pressure.

When similar measurements are made using the Fahrenheit scale of temperature, absolute zero is found at *-459.67 °F.* When this point is used as the zero and Fahrenheit degrees are used, the scale is called the *Rankine scale.* The chart shows the various temperature scales and their relations:



CONVERSIONS. . . . .

The simple rules below will help you convert from one scale to another:

To Convert From	To	Add / Subtract / Multiply By. . .
°C or °F	°K or °R	Add 273° or 460°
°K or °R	°C or °F	Subtract 273° or 460°
°K or °R	°R or °K	Multiply By. . . 9/5 or 5/9

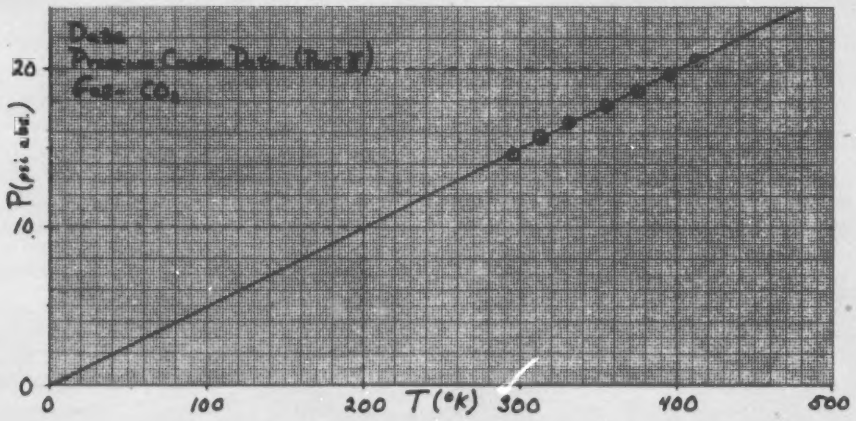
KELVIN TEMPERATURE SIMPLIFIES THE P-T RELATION....

The reason for introducing absolute zero and Kelvin temperature is to simplify our analysis of the P-T relation of a gas and thus get us closer to the mathematical expression for the General Gas Law. To see this do the following:

6. REPLOT YOUR DATA IN KELVIN TEMPERATURE:

From what we have said you can now replot your data in terms of Kelvin temperature. First convert your temperature data from °C to °K and then make a new graph of P vs T.

The graph now shows our sample data plotted as pressure against *Kelvin* temperature. Notice that the pressure is now zero when the temperature is zero. Also note that when the temperature is doubled (from say 200°K to 400°K) the pressure doubles from 10 psi to 20 psi). This means that *the absolute pressure and absolute (Kelvin) temperature of a gas are directly proportional.*



As a mathematical expression we can thus write:

$$P \propto T$$

Absolute Gas Pressure — is Proportional to — Absolute Gas Temperature

or its equivalent: $P=CT$

This then is the P-T relation for a gas when the volume and number of molecules is held constant. Though we have not shown it, this relation holds true for all gases, regardless of their type. In your case the gas was air, a mixture of nitrogen and oxygen, but your results would have been the same for the two gases separately or for any other gas you might choose, helium, CO_2, steam, etc.

OPTIONAL EXPERIMENT:

To verify the above statements you may wish to redo this experiment using another gas. To compare the results, extrapolate your data to 0 pressure and see that the value of absolute zero is the same. Your value for C will compare only if your initial conditions of P and T are identical! The reasons for this will become clear as we continue.

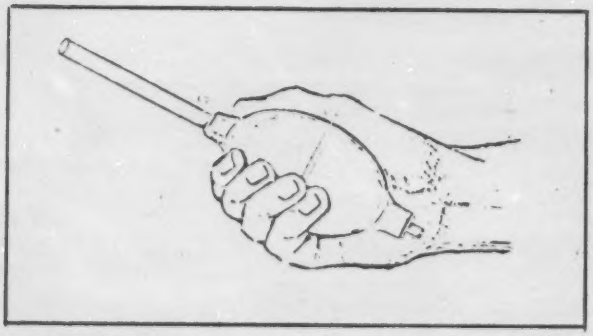
OTHER P,V,T,N RELATIONS--FROM YOUR EXPERIENCE.....

The pressure cooker experiment has thus given us the P-T relation for a gas. But there are two more parameters, volume, V, and number of molecules, N, and many more relations to determine before a complete expression for the General Gas Law can be obtained. Many of these relations are at least qualitatively familiar to you from your own experience. We will discuss some of these here and also give the relation that would result if quantitative experiments were performed. By summarizing these relations we will then be able to suggest a form for the General Gas Law which will summarize them all.

Volume:

Volume is a parameter with which you should now be quite familiar. It is generally measured in cubic measure; cubic inches (in^3), cubic centimeters (cm^3), etc. However, in the case of the pressure cooker it is specified in quarts, a common volume measurement for liquids. The conversions between several units are given opposite. Note that 1 quart is about equal to 1 liter which is exactly 1000 cm^3. This is a useful rough conversion between the English and metric systems.

	Quarts	in^3	Liters	cm^3
1 Quart =	1	57.7	.946	946
1 in^3 =	.0173	1	.0164	16.4
1 Liter =	1.06	61.0	1	1000
1 cm^3 =	.00106	.0610	.00100	1



V - P RELATION: (Boyle's Law)

The volume of a gas is of course related to its pressure. A simple way to convince yourself of this is to recall the bulb we used to pump up the pressure cooker. If you close the air release hole (thus fixing the number of molecules) and squeeze the bulb, the volume of air in the bulb decreases but the force you must exert gets larger. The fact that you must squeeze harder indicates that the air pressure is increasing. Thus you see that volume and pressure are inversely related. If P is the *absolute* gas pressure, then P and the volume V are related very nearly by:

$$V \propto \frac{1}{P}$$

V - T RELATION: (Gay-Lussac's Law)

If we again work with a definite number of molecules, but now a definite pressure (fix N and P), then we find that volume increases as temperature increases. This effect leads to the familiar observation that "hot air rises and cold air sinks". When air is heated it expands increasing its volume which makes it lighter. This lighter air then tends to float or rise above colder, heaver air. This is the principle behind convection. If *Kelvin* temperature is used the expression describing this phenomenon is:

$$V \propto T$$

Number of Molecules:

Earlier we said that N was the number of molecules of the gas and the General Gas Law was concerned with N. We should mention, however, that chemists would rather speak about the number of moles (n) of a gas and you will find that the General Gas Law is often written in terms of number of moles. The two, of course, are related. *There are 6.02252 x 10²³* molecules in a mole of a gas. (Avogadro's number) The reason for prefering moles is obvious it's much easier to say 3 moles than 18 x 10^{23} molecules. But physicists prefer thinking in terms of molecules and that is what we'll use. But it's important to remember that here we are concerned only with the number of gas molecules, *not* the mass of gas, and it is this difference that makes the law general.

PHYSICS / CHEMISTRY CONVERSION

Molecules = # Moles x Avogadro's

$$N = n \times 6.02 \times 10^{23}$$

But, how do you get the value of N for a given quantity of gas. Clearly you cannot count up the number of molecules. However, you *can* measure its mass. *Mass* and *number of molecules* are simply related through the *atomic number* of the gas involved. That is if a certain type of gas has an atomic number, A, then there are A grams of that gas in a mole. And we know that there are 6.02 x 10^{23} molecules in a mole of any gas. Multiplying this number times the mass of the gas you have will then give the number of molecules you have. Of course if it is a mixture of gases, like air, then you must know the fraction of each and then calculate the number of each kind of molecule using its value for A,.....a rather complicated process.

MOLECULES / MASS CONVERSION

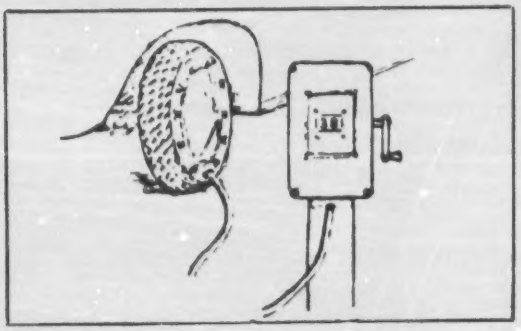
$$\#\ \text{Mol.} = \text{Mass(gms)} \times \frac{\text{Avogadro's\#}\ \text{(mol./mole)}}{\text{Atomic \#}\quad\ \text{(gms/mole)}}$$

$$N = M \times \frac{6.02 \times 10^{23}}{A}$$

N - V RELATION:

The relation between N and V can be seen by considering a paper bag. When you blow up a paper bag you are simply putting more molecules into the bag. The temperature of the bag does not rise (except that your breath is warm) and the pressure inside doesn't increase (T and P fixed), the bag just opens up more to increase its volume.

The quantitative description of this is simply: **VαN**



N - P RELATION:

The N-P relation can be illustrated by considering the tires in your car. When you inflate your tires, the service station pump forces air molecules into the tire, the tire volume and temperature stay essentially the same but a pressure gauge shows that the pressure has increased. You saw this same phenomenon when you pumped up the pressure cooker to leak test it. If *absolute* pressure is used, then the N-P relation becomes:

57 **PαN**

THE GENERAL GAS LAW...

SUMMARIZES THE P, V, N, T RELATIONS. . .

All of the relations we have discussed may be summarized by the proportion:

$$PV \propto NT$$

Compare this expression with the results of previous pages to convince yourself that it agrees with all of the earlier discussions!

P	$\propto$	T
P	$\propto$	N
V	$\propto$	1/P
V	$\propto$	T
V	$\propto$	N

If we now define a constant of proportionality k we will have the mathematical expression of the General Gas Law:

$$PV = NkT$$

This single equation expresses all that we have said so far about the behavior of gases. As a mathematical relation it can be used to calculate changes in one of the parameters when others are changed. It must be remembered however, that *absolute values for pressure and temperature must be used* in all such calculations.

k, Boltzmann's Constant Of Proportionality:

The constant of proportionality, k, is called *Boltzmann's Constant* after Ludwig von Boltzmann. It is one of the fundamental physical constants and has the value 1.38×10^{-23} joules/°K in MKS units. Its physical significance will be discussed in the following pages.

Consistent Units:

As in all mathematical expressions one must use the same type of units for all the parameters, for example: all MKS, all cgs,etc. The table below gives these various units:

	MKS	cgs	English
P	newtons/m^2	dynes/cm^2	lbs/in^2
V	m^3	cm^3	in^3
T	°K	°K	°R
N	# molecules	# molecules	# molecules
k	1.38×10^{-23} joules/°K	1.38×10^{-16} erg/°K	6.79×10^{-23} in.lbs/°R

. . .AND IS USEFUL:

Now that we have an expression for the behavior of a gas let us see how it can be used. Following are two calculations that demonstrate the type of questions that the General Gas Law can help answer. In these examples it is not necessary to use the whole gas law expression but simply to recognize which parameters change and which are fixed, and then to apply the proper proportionality relation.

HOW MUCH DOES TIRE PRESSURE INCREASE AS TIRES HEAT UP?

Assume an automobile tire was inflated to 30 psi (gauge) when the temperature was 60°F. After some high speed driving the tire temperature rises to 150°F. What is the pressure of the air in the tires now?

REPRINTED FROM AN AUTOMOBILE OWNERS MANUAL →

Tire tread life varies from car to car because of driver habits. Aside from good driving practices, the most important factor in obtaining maximum tread life is maintaining proper tire pressure. Pressure lower than recommended will reduce the allowable full rated load and may affect vehicle handling. Over-inflated tires reduce ride comfort by magnifying rather than absorbing road shocks and are more vulnerable to damage from road surface impacts.

Check tire pressure frequently and while "cold" (preferably after the car has been parked at least one hour and before driving more than three miles).

Operation at high speeds or heavy loads will increase tire pressure considerably. An 8 pound increase over "cold" pressure is not unusual. Therefore never bleed air from a "hot" tire because the pressure is above normal.

From the statement of the question and our own experience we can deduce the following:

Initial Tire Conditions: $P_i = 30$ psi (gauge), $T_i = 60°F$

Final Tire Conditions: P_f = to be determined, $T_f = 150°F$

We can assume that the tire does not leak so that: $N_i = N_f = N$ fixed

.....and that the tire doesn't increase significantly in size, or: $V_i = V_f = V$ fixed

Writing the General Gas Law for these two conditions;

$$P_f V = NkT_f$$

$$P_i V = NkT_i$$

..and dividing one by the other gives the simple proportionality:

$$\frac{P_f}{P_i} = \frac{T_f}{T_i}$$

Since the General Gas Law requires absolute pressure and absolute (Kelvin) temperature we must first convert our values of P_i, T_i and T_f to these:

$P_i = 30$ psi (gauge) $= 45$ psi (abs.)
$T_i = 60°F = 16°C = 289°K$
$T_f = 150°F = 65°C = 338°K$

Substituting this into the equation gives:

$$\frac{P_f}{45} = \frac{338}{289} \quad \text{or} \quad P_f = \frac{338}{289} \times 45$$

or $P_f = 53$ psi (abs.)

59

or $P_f = 38$ psi (gauge)

HOW HIGH DOES THE PRESSURE GET IN AN ENGINE CYLINDER?

*The compression ratio of a certain gasoline engine is 9.5:1. That is, the volume of gas in the cylinder when the piston is at the bottom of its stroke is 9.5 times the volume at the top of its stroke. If a charge of fuel-air mixture is drawn in at the bottom of the stroke at 2.0 psi vacuum what is the pressure in the piston at the top of the stroke? Assume the process to be isothermal, that is that the temperature of the mixture does not change.**

REPRINTED FROM AN AUTO REPAIR MANUAL →

USING A COMPRESSION GAUGE

Another very important engine testing gauge is the compression tester. It measures the pressure within the cylinder in pounds per square inch (psi). As with the vacuum gauge, the theoretical numerical reading is not so important as the variation between cylinders. The cylinder pressures should not vary over 15 psi; otherwise, the engine cannot be tuned properly. Variations cause uneven idling and loss of power.

To use the gauge, remove all the spark plugs and insert the rubber tip into each spark plug hole in turn. With the throttle held wide open, crank the engine to obtain about 6 power impulses on the gauge; record the reading. Do this at each cylinder and compare the results. Generally, modern high-compression engines have a reading close to 175 psi.

From the statement of the question we can write down that:

Initial Fuel conditions: P_i = 2.0 psi vacuum; V_i = not known but = $9.5V_f$

Final fuel conditions: P_f = to be determined; V_f = not known

The assumption of an isothermal compression means that; $T_i = T_f = T$ fixed

. . . . . and we assume there is no loss of mixture so that: $N_i = N_f = N$ fixed

Writing the General Gas Law for these two conditions gives:

$$P_iV_i = N_ikT_i$$

$$P_fV_f = N_fkT_f$$

Substituting the symbols that we know and dividing one equation by the other gives:

$$\frac{P_f \quad V_f}{P_i \quad (9.5V_f)} = \frac{NkT}{NkT}$$

or the proportionality:

$$\frac{P_f}{9.5P_i} = 1$$

or:

$$P_f = 9.5P_i$$

Recalling that the General Gas Law requires absolute values: P_i = 2 psi vacuum = 12.7 psi abs.

Substituting this into the equation gives: P_f = 9.5 X 12.7 psi abs.

or: P_f = 121 psi abs.

or: P_f = 106 psi gauge

* In actual fact the temperature rises considerably since work is done on the gas as it is compressed, and the resulting thermal energy does not have time to get dissipated to the cooling system during the compression stroke. This rise in temperature means a corresponding rise in pressure according to the General Gas Law. Thus the calculation can only give a lower limit to the pressure produced in the cylinder. Typical real data is given in problem 10 at the end of this section.

A MOLECULAR MODEL FOR THE GAS (Optional):

SCIENTIFIC MODELS - WHAT AND WHY?

In this part of the module we have obtained an expression for the General Gas Law based on our observation of the way in which gases behave. It is expressed in term of quantities which we can measure; pressure, temperature, volume, Formulas derived in this way, that is from observations only, are called *empirical formulas*, and are very often the first step in the scientific understanding of a phenomenon. One of the goals of physics however, is to understand why the empirical formula is what it is in the hopes that knowing the basis of the formula will lead to additional formulas which we did not measure. In order to do this a scientist usually thinks up a theoretical model for the phenomenon; that is he describes a certain picture of what is happening. From this model he then tries to derive by calculation the original empirical formula and any other ones that he can, revising his model as needed to better fit the experimental observations.

These derivations and calculations can become quite complicated and sophisticated. The models on which they are based, on the other hand, may be fairly simple and very often can provide a useful way of thinking of the phenomenon and how things will behave.

In the next few pages we will describe briefly the model for a gas which leads to the General Gas Law without giving you the detailed derivations. From the model and the Law we can then understand a bit about why the gas behaves as it does and obtain formulas for the specific heat of the gas to compare with what we found in part one.

THE MODEL FOR AN "IDEAL" GAS:

If one were able to take an instantaneous magnified picture of the cross-section of the pressure cooker wall one might see something like that shown in the diagram on the next page. The aluminum atoms of the pressure cooker wall are all tightly packed together. They do not stray from their positions but simply vibrate slightly about their patterned positions. The molecules of the air on the other hand are free running. They bang about at great speeds like a bunch of bb's in a can, bouncing off each other and the walls many times a second without sticking or getting significantly slowed down. As can be seen they are quite sparse in number compared to those of the solid aluminum walls. This of course accounts for the fact that a volume of gas is much lighter than an equal volume of solid.

This is the model we use for a gas and which will lead to the General Gas Law. The basic assumptions which this model makes about the gas are the following:

1. All molecules, regardless of what kind, may be treated the same way but with their own mass m.

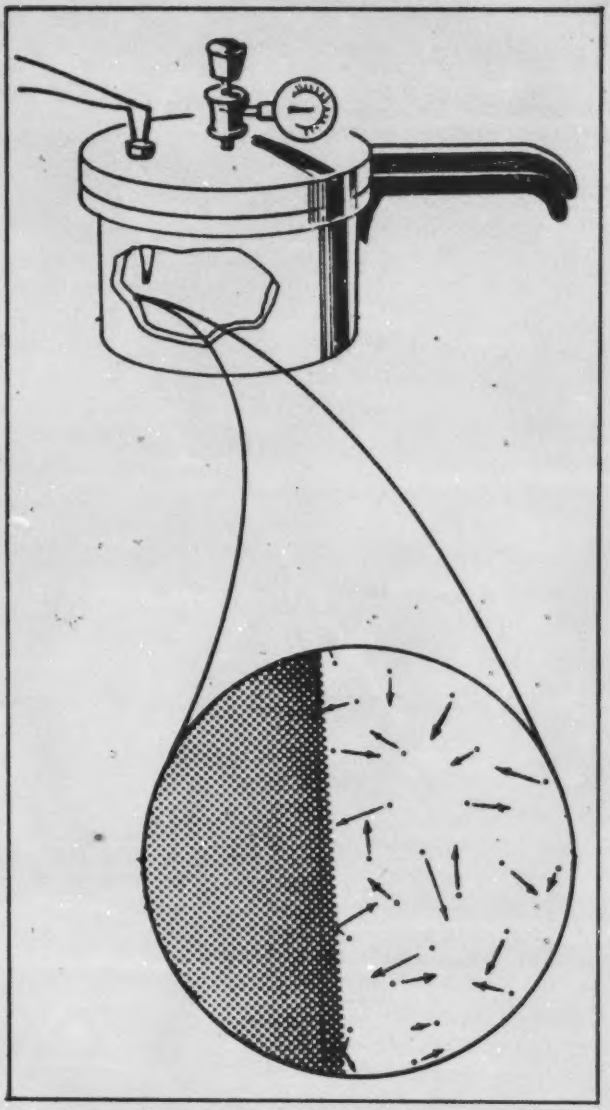
2. The size of molecules are small compared to their separation; they are essentially a bunch of small points of mass m.

3. When two molecules strike each other, they both act as hard spheres, bouncing off without sticking and without losing any energy.

4. The energy of the gas as a whole is made up simply of the kinetic energy of motion of the various molecules.

Real gases don't always behave this way so that these assumptions are those that apply only to what we called an "ideal" gas. However the "ideal" gas model works quite well in many cases and theoretical calculations based on it explain the General Gas Law and many other properties of the gas as we will see.

According to the theoretical model for an "ideal" gas, a microscopic photograph of the inside of the pressure cooker would reveal the gas to be a sparse collection of minature bb's bouncing around in all directions at great speeds.



WHAT CAN THE GENERAL GAS LAW TELL US ABOUT THE MODEL?

The detailed calculations of the theoretical model entail rather sophisticated mathematical techniques so that we will not go into them here. However, by looking closely at the General Gas Law we can get a hint of several things about the model that can be quite useful.

How Far Apart Are The Molecules In A Gas?

One thing we can ask is the average space that a molecule has to move around in. That is the volume per molecule, V/N. According to the General Gas Law this is:

Volume per gas molecule $$\frac{V}{N} = k\frac{T}{P}$$

We did not measure the volume of our sample of gas, nor do we know the number of molecules. But the equation says we can calculate V/N from the known constant k, and any two values of T and P, *measured at the same time.*

Let us use the sample data shown on the graph of page We can select values from any pair on the line, for example at a temperature of 100°C the pressure is 18.7 psi abs.

First we must convert the temperature to °K; $T = 100°C + 273 = 373°K$

...and then the pressure to nt/m^2: $$P = 18.7 \text{ psi} \times 6.89 \times 10^3 \frac{nt/m^2}{\text{psi}}$$

$$= 1.28 \times 10^5 \text{ nt/m}^2$$

Putting these and Boltzmann's constant into the equation gives:

$$\frac{V}{N} = 1.38 \times 10^{-23} \frac{373}{1.28 \times 10^{5}}$$

$$\frac{V}{N} = 4.0 \times 10^{-26} \frac{m^3}{\text{molecule}}$$

The average volume allowed for each air molecule:

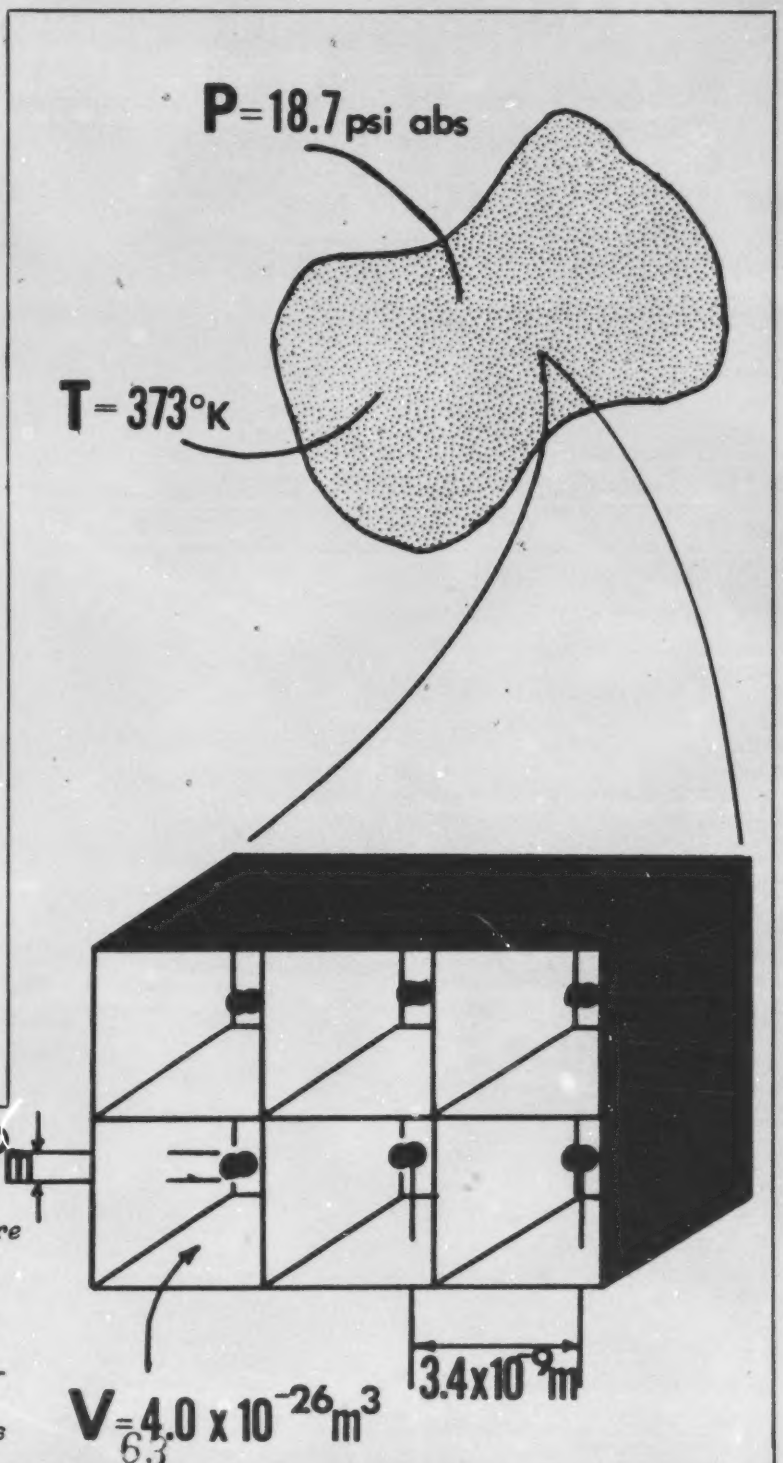
Thus on the average, for every 4.0×10^{-26} cubic meters of space there is one molecule. If we were to divide the total space into cubes of this volume, as shown, and put one molecule in the center of each cube then the average separation of the molecules would be the same as the length of the side of the cube. Thus the cube root of 4.0×10^{-26} cubic meters:

$$S = \sqrt[3]{4.0 \times 10^{-26}}$$

would give: ***The average separation of air molecules:***

$$S = 3.4 \times 10^{-9} \text{ meters}$$

It is known from other measurements that a nitrogen molecule has a diameter of around 4×10^{-10} meters. Thus the average molecular separation is about 10 times the molecular diameter and assumption 2 for the ideal gas was a good one.



If the pressure and temperature of a gas are 18.7 psi abs and 373°K, then there is one atom for every 4.0×10^{-26} m^3 of space. This means that on the average the molecules are separated by a distance of 3.4×10^{-9} m which is about 10 times their diameter.

How Much Energy Does A Gas Molecule Have?

Now let us look at the product PV. What are its units?

in MKS units:

PV = Pressure x volume
= Force/Area x volume
= Newtons/m^2 x m^3
= Newtons-meters

Newton-meters are joules, units of energy. If the product PV is an energy then so must NkT be an energy since the General Gas Law says PV = NkT. Thus the General Gas Law expresses an energy relation.

If we divide PV by the number of gas molecules present, N, then PV/N would seem to express the energy per molecule.

That is, the relation;

$$\frac{PV}{N} = kT$$

suggests that:

The energy per molecule | is proportional | the Kelvin temperature

By using a more sophisticated mathematical analysis it can be shown that the *average kinetic energy* of a gas molecule, $1/2mv^2$, is given exactly by:

Average kinetic energy per gas molecule:	$1/2\ mv^2 = 3/2\ kT$

were m is the mass of a single molecule and v is its *average* speed. (Actually v is a special kind of average, called the root mean square value, but for our purposes the difference is not important).

Thus we see that the average kinetic energy of a gas molecule depends only on the absolute temperature of the gas and nothing else.

WHAT CAN THE MODEL TELL US ABOUT THE GAS?

What Is The Total Energy Of The Gas?

According to assumption 4 of our model, the energy of the gas is made up solely of the kinetic energy of motion of the individual molecules that compose it. We have just seen that the average kinetic energy of each molecule is 3/2 kT. Thus if there are N molecules in a quantity of gas then there must be N times 3/2kT total energy.

Total energy in the gas:	$E = 3/2\ NkT$

According to our "ideal" gas model, then, the total energy of the gas depends only on its temperature and nothing else.

What Is The Specific Heat Of The Gas At Constant Volume?

Now that we have an expression for the total energy we can quite simply determine the amount of energy we would have to add to change the temperature by one degree. That is if we have a closed container of gas having N molecules;

its energy is: $E_1 = 3/2\ NkT_1$

If we raise its temperature by 1°, $T_2 = T_1 + 1°$

without letting it change its volume

then the new energy is:

$$E_2 = 3/2\ Nk\ (T_1 + 1°)$$
$$= 3/2\ NkT_1 + 3/2\ Nk$$

and we have added an amount of energy:

$$Q_v = E_2 - E_1$$
$$Q_v = 3/2\ Nk$$

This quantity is very similar to our earlier definition of specific heat. If N is the number of molecules in a unit mass of a substance then this is the amount of energy required to heat a unit mass of a substance by one degree when the volume is kept constant. Recall that the value of N for a gas of molecules with atomic mass A and mass M is:

$$N = 6.02 \times 10^{23}\ \frac{M}{A}$$

Our earlier definition of specific heat was: $C_v = \dfrac{Q_v}{M\Delta T}$

If we substitute in the value of N in Q_v and let $\Delta T = 1°$ we have; $C_v = \dfrac{3/2 \times 6.02 \times 10^{13}\ Mk}{M \times 1}$

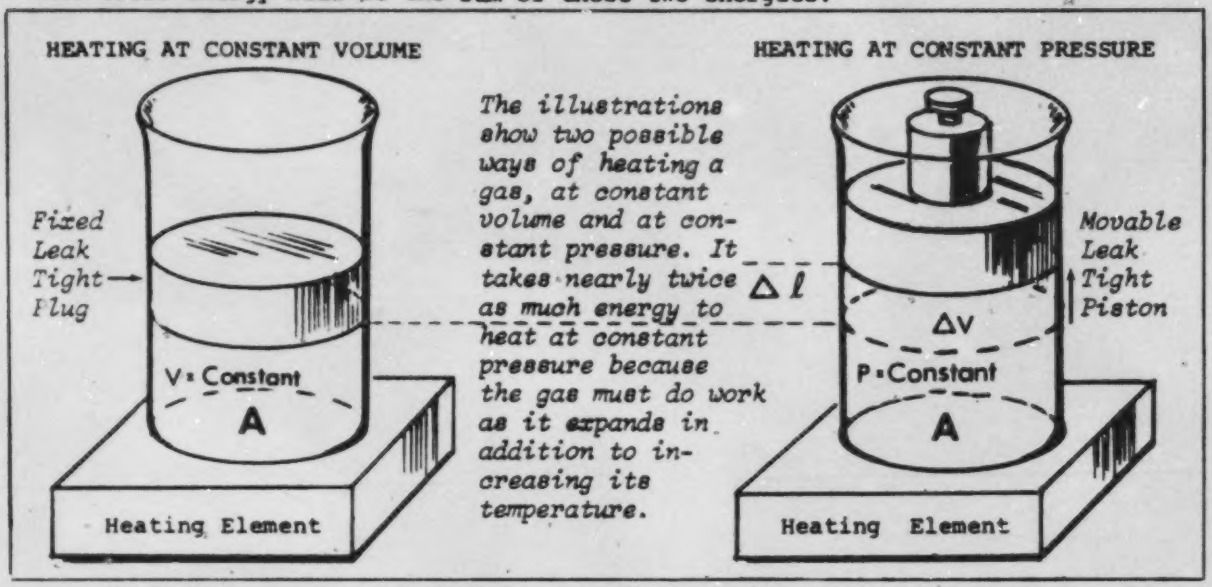
The specific heat at constant volume:	$C_v = 3/2 \times 6.02 \times 10^{23}\ \frac{k}{A}$

According to our model, the specific heat of gases depends only on their atomic mass A and nothing else. *The bigger the atomic number the smaller the the specific heat.*

What Is The Specific Heat Of A Gas At Constant Pressure?

In the derivation above we were very careful to say that we raised the temperature *"at constant volume"*. One might ask would their be any difference if we fixed the pressure for the process rather than the volume. The illustrations on the next page show the two possibilities. Note that for the heating

at constant pressure, the gas must expand in order to keep the pressure the same. In doing so it exerts a pressure force against the container walls moving them outward. If we try to harness this effort by letting it move a piston, then it appears as work which we can calculate. Thus the energy we put in must not only heat up the gas (as it did at constant volume) but also do work as it expands. The total energy will be the sum of these two energies.


HEATING AT CONSTANT VOLUME
HEATING AT CONSTANT PRESSURE
Fixed Leak Tight Plug
V = Constant
A
Heating Element
The illustrations show two possible ways of heating a gas, at constant volume and at constant pressure. It takes nearly twice as much energy to heat at constant pressure because the gas must do work as it expands in addition to increasing its temperature.
Δ l
ΔV
P = Constant
A
Heating Element
Movable Leak Tight Piston

To calculate the work done in expanding we will use the illustrated example. The work done in moving the piston a distance $\Delta\ell$ is:

$$W = F\Delta\ell$$

The pressure force on the piston is simply the pressure in the cylinder times the piston area A;

$$F = PA$$

so that the work becomes:

$$W = PA\Delta\ell$$

The product $A\Delta\ell$ is the increase in cylinder volume ΔV;

$$\Delta V = A\Delta\ell$$

so that the work becomes:

$$W = P\Delta V$$

If we write the General Gas Law for the initial;

$$PV = NkT$$

and final condition we have:

$$P(V + \Delta V) = Nk\ (T + \Delta T)$$

Rewriting and subtracting gives:

$$PV + P\Delta V = NkT + Nk\Delta T$$

$$\underline{PV = NkT}$$

$$P\Delta V = Nk\Delta T$$

Substituting this into the expression for work gives:

$$W = Nk\Delta T$$

The amount of work done for a 1° increase in temperature ($\Delta T = 1°$) gives:

$$W = Nk$$

Thus the total energy required to heat the gas *at constant pressure* is:

$$Q_p = Q_v + W$$
$$Q_p = 3/2Nk + Nk$$
$$Q_p = 5/2Nk$$

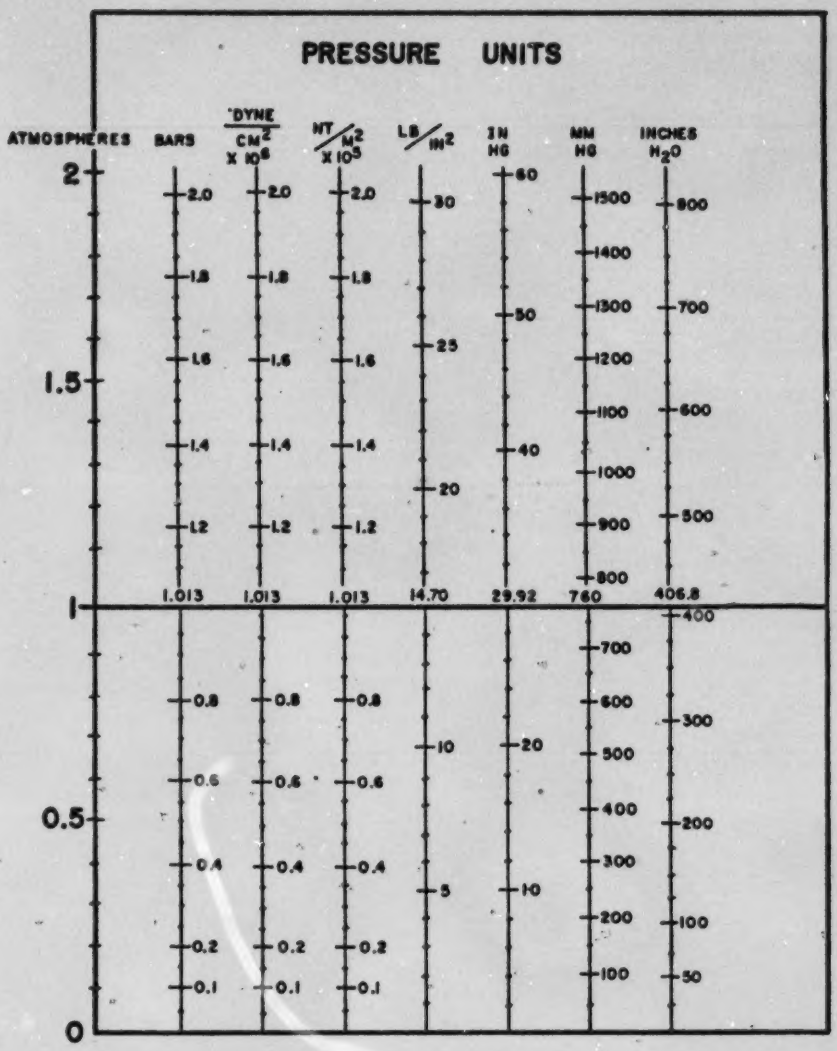
We see that it takes almost twice as much energy to heat a gas at constant pressure as it does to heat it at constant volume.

We can calculate as before to get:

The specific heat at constant pressure:	$C_p = 5/2 \times 6.02 \times 10^{23} \frac{k}{A}$

Again according to our model as the atomic mass goes up the specific heat goes down. Refer again to the table of specific heats of various gases found on page . These gases are arranged in order of increasing atomic mass. Note that except for CO_2, our model has given the right results. Exceptions like CO_2 indicate that our model was too simple and other assumptions must be made.

PRESSURE UNITS....



There are many units for measuring pressure in common use. The chart on the right shows some of them and how they are related. You are likely to encounter any one of them and very often more than one at the same time as in the gauge you used to measure the pressure in the cooker. There are three broad methods for specifying pressure.

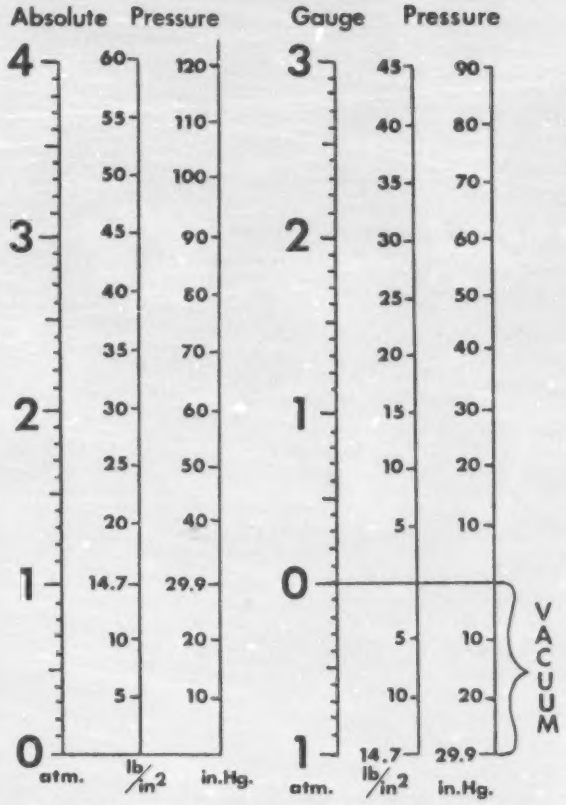
FORCE PER UNIT AREA: One is in terms of force per unit area as we discussed on the previous pages. In the English system this is *pounds per square inch*, lbs/in^2 (or psi) and in the metric system, *newtons per square meter* (nt/m^2) or *dynes per square centimeter* (dyne/cm^2). As you can see from the graph one atmosphere (or atmospheric pressure) is about 14.7 psi, 1.013 x 10^5 nt/m^2 or 1.013 x 10^6 dynes/cm^2.

ATMOSPHERES: Sometimes, particularly when specifying very high pressures, it is more convenient to use simply the normal or average atmospheric pressure as the unit. This unit is called the *atmosphere* (atm). A nearly equivalent unit is the *bar*. The bar is a convenient unit since it is nearly equal to one atmosphere yet is exactly 10^5 nt/m^2 in the metric system.

LIQUID COLUMNS: In meteorology (and many other scientific applications) pressures are measured in terms of the height of the column of liquid that the pressure will support. Atmospheric pressure will support a mercury column 29.92 inches or 760 mm high. Therefore pressure are often expressed in inches or millimeters of mercury (in. Hg or mm Hg). Your vacuum readings in this experiment are expressed in inches of mercury below atmospheric. Note that atmospheric pressure will support a column of water 34 feet or about 407 inches high. Inches of water is a convenient unit for measuring low pressure.

ABSOLUTE PRESSURE -- GAUGE PRESSURE -- VACUUM. . . .

Of course, even though our atmosphere is not in a closed container, it exerts a pressure on any surface in the atmosphere. All over our bodies we feel a pressure of about 14.7 lbs/in^2 - the normal atmospheric pressure at sea level.



ABSOLUTE: There are two ways of stating pressure. *If you measure it starting from zero pressure, then it is called absolute pressure.* When we say atmospheric pressure is 14.7 psi, this is absolute pressure. The pressures shown on the previous page are absolute pressures. Very often an "a" is put after the pressure unit to indicate absolute pressure; for example psia.

GAUGE: However, since atmospheric pressure is always with us it is often more useful to specify the pressure relative to atmospheric pressure. That is, we pretend that atmospheric pressure is 0 (as you saw on your pressure gauge) and specify the pressure relative to that. *The pressure measured relative to one atmosphere is called gauge pressure.* When you inflate your tires to 30 lb/in^2, this is 30 psi gauge pressure. The letter "g" is often put after the unit to indicate gauge pressure; for example psig. To get the absolute pressure in the tires you must add 14.7 lb/in^2 absolute pressure.

$$P\ (abs) = P\ (gauge) + P\ (atmospheric)$$

VACUUM: *Gauge pressure below atmospheric is called vacuum,* since one has to evacuate a region to get below atmospheric. A perfect vacuum exists when there is 0 absolute pressure. The units specifying vacuum are then the reverse of the units from 0 to one atmosphere. That is a perfect vacuum (gauge pressure) is 14.7 psi or 29.92 in/Hg, 760 mm Hg, etc.

$$P\ (absolute) = P\ (atmospheric) - P\ (gauge\ vacuum)$$

Pressure Units and Conversion Factors

A \ B	atm	bars	$\frac{dyne}{cm^2}$	$\frac{nt}{m^2}$	$\frac{lb}{in^2}$	in Hg	mm Hg	in H_2O	Unit Systems
atm	1	1.013	1.01×10^{6}	1.013	14.7	29.9	760	407	
bars	.987	1	10^{6}	.987	14.5	29.5	750	402	MKS
$\frac{dyne}{cm^2}$	9.87×10^{-7}	10^{-6}	1	.100	1.45×10^{-5}	2.95×10^{-5}	7.50×10^{-4}	4.02×10^{-4}	cgs
$\frac{nt}{in^2}$	.987	1.013	10	1	1.45×10^{-4}	2.95×10^{-4}	7.50×10^{-3}	4.02×10^{-3}	The SI Unit MKS
$\frac{lb}{in^2}$	.0680	0.0689	6.89×10^{4}	6.89×10^{3}	1	2.04	51.7	27.7	English
in Hg	.0334	.0338	3.39×10^{4}	3.38×10^{3}	.491	1	25.4	13.6	English
mm Hg (Torr)	1.32×10^{-3}	1.33×10^{-3}	1330	133	1.94×10^{-2}	.393	1	.535	cgs
in H_2O	2.46×10^{-3}	2.49×10^{-3}	2490	249	.0361	.0736	1.87	1	English

HOW TO USE THE TABLE:

If you have a unit in column A and you want a unit in Column B multiply by the factor shown.

FOR EXAMPLE: *If you have* $\underline{lb/in^2}$ *and you want* $\underline{mm\ Hg}$ *multiply by* 51.7 $\frac{mm\ Hg}{lb/in^2}$

$$A(lb/in^2) \times 51.7\ \frac{mm\ Hg}{lb/in^2} = B(mm\ Hg)$$

Review

SUMMARY

Part II of the module is concerned with the behavior of gases. We usually describe the condition or *state* of a gas by listing four quantities - called *variables* or *parameters: pressure*,P; *volume*,V; *number of molecules*,N; *temperature*,T. Changes in the state of a gas are described by the *General Gas Law*.

In the experiment a sealed pressure cooker (N and V held constant) is used to measure changes in gas pressure as the temperature changes.

Pressure is defined as force per unit area - P = F/A. When pressure is acting over an area, the total force is given by F = PA. *Absolute* pressure means pressure measured from zero. *Atmospheric* pressure at sea level is about 14.7 lb/in^2 absolute. *Gauge* pressure is the pressure measured *above* atmospheric pressure. *Vacuum* is the amount of a pressure *below* atmospheric.

The experiment showed: pressure is proportional to temperature. *Extrapolation* of the P-T graph indicates that the pressure of a gas would become zero at -273°C. Other evidence makes us regard this as the lowest possible temperature - *absolute zero*. There are two *absolute temperature* scales: *Kelvin* and *Rankine*. °K = °C + 273°.

The results of the experiment, plus other observations about the behavior of gases can all be summed up in one equation, the General Gas Law:

$$PV = NkT$$

The constant k is called *Boltzmann's constant*.

PROBLEMS

1. SCUBA tanks are generally filled with compressed air at 2200 psi absolute. A certain tank was filled to this pressure at a temperature of 20°C (68°F). Later this tank was left for several hours in the trunk of an automobile, where the temperature reached 66°C (150°F). The bursting pressure of the tank is rated at 2900 psi. Do you consider this a safe thing to do?

   Explain.

2. A welder's tank of acetylene gas has a volume of 0.1 m^3. At a temperature of 20°C, the pressure of the gas is 9 atmospheres gauge. How many acetylene molecules are in the tank?

3. An air bubble of 15 cm^3 volume is at the bottom of a lake where the temperature is 4°C and the pressure 4 atmospheres gauge. The bubbles rises to the surface where the water temperature is 20°C. Assume that the air in the bubble has the same temperature as the surrounding water; calculate the volume of the bubble just before it reaches the surface.

4. Oxygen boils at -182.97°C. Express this temperature on the Kelvin scale.

   Zinc melts at a temperature of 419°C. Express this on the Fahrenheit and Rankine scale.

5. The normal human body temperature is 98.6°F. Express this on the Celsius and Kelvin scales.

   The temperature of the sun's surface is about 6000°K. Express this in degrees Fahrenheit.

6. A certain compound pressure gauge reads 12.5 in Hg vacuum. What is the absolute pressure in inches of Hg? Express this pressure in atmospheres. Express this pressure in newtons per square meter.

7. The pressure in sea water increases one atmosphere for every 33 ft. of depth. What is the gauge pressure at a depth of 132 ft.? The absolute pressure? (Express in psi).

   What would be the total force on a surface of 10 in^2?

8. The pressure inside a laboratory autoclave (similar to a pressure cooker) is to be regulated to a maximum value of 30 psig. The regulating weight rests over a tube of area 0.05 in^2? What should be the weight of the regulating weight?

9. The piston of an air compressor has an area of 15 in^2. At one point in its stroke the total force exerted by the piston is 2500 pounds. What is the pressure of the air contained in the cylinder of the compressor?

10. The compression ratio of a certain gasoline engine is 9.5:1. That is, the volume of gas in the cylinder when the piston is at the bottom of its stroke is 9.5 times the volume at the other end of the stroke.

A certain charge of fuel-air mixture at the maximum volume has a pressure of 12 lb/in^2 at a temperature of 70°C. When fully compressed in the cylinder the temperature rises to 300°C.

What is the pressure of the fuel-air mixture when it is fully compressed in the cylinder?

BOILING TEMPERATURE AND PRESSURE

INTRODUCTION

In part one you studied changes of phase.

You supplied energy at a constant rate to a known mass of ice and measured changes in temperature. First the ice warmed up to 0°C, then it melted - that was a phase change. Then the water steadiiy warmed up to 100°C, then it began boiling. The temperature remained constant as the water boiled - another phase change. All this happened at a constant pressure of one atmosphere.

In part two you studied the gas phase.

By heating air in the pressure cooker you saw how the pressure and temperature of a gas are related. The General Gas Law ($PV = NkT$) describes the behavior of the gas phase in terms of a simple equation.

In part three you will see how phase changes are related to pressure and temperature.

In the experiment you will see how the boiling point is affected by pressure. These phase changes turn out to be related to many aspects of daily life - ranging from weather to gas storage.

THE RELATION BETWEEN PRESSURE AND BOILING TEMPERATURE. . . . .

The fact that the boiling point changes with pressure means that one has some control over the temperature at which liquids boil, or equivalently, the temperature at which gases condense. Since high pressures can be dangerous there is a practical limit to what one can do, but in the pressure cooker and in the examples below a moderate pressure difference has significant repercussions.

. . . . .HAS MANY TECHNICAL APPLICATIONS:

SOME GASES MAY BE STORED AT HIGH PRESSURES AS LIQUIDS

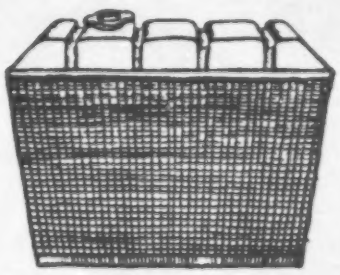
Natural gas and propane are commonly used for cooking and heating in thousands of homes and buildings. As gases they are rather difficult to store efficiently since they take up so much room. Thus for house trailers, campers, torches, and even automobiles they would be totally inappropriate fuels unless they could be condensed as liquids. At one atmosphere you would have to cool propane to below -40°F to make it condense to a liquid. But if propane is pressurized to around 200 psi, it condenses to a liquid and remains liquid until the temperature reaches 100°F or more. As a liquid, propane takes up much less space so a large quantity can be stored in a rather small container. Thus, "bottled gas" is really not gas at all but rather "bottled liquid" that converts to a gas when it is released to one atmosphere.

"Bottled Gas" is conveniently stored as a high-pressure liquid.

COOLING SYSTEMS OPERATE MORE EFFICIENTLY AT HIGH TEMPERATURES

The radiator of your car is sealed with a cap that allows the gauge pressure in the cooling system to build up to 12 to 15 psig. With water in the system the temperature could rise to 245°F without boiling. Most manufacturers, ("Prestone" and others), now recommend a cooling liquid that is 50% ethylene glycol with water. This mixture does not begin to boil until it reaches a temperature of 260°F or higher.

With the cooling liquid at a high temperature (but not boiling) the radiator cools more efficiently, since the temperature difference is large. The car-maker, therefore, can use a smaller radiator and a smaller opening in front for air flow. You get better visibility and a sportier-looking front end.



Car radiators use high boiling point liquids to make them more efficient.

DANGER: *There is one big danger, however. If you remove the cap, you suddenly release the pressure and the coolant will suddenly boil violently. The reason for this will become clear as you read this section.*

PHASE DIAGRAMS

A WAY OF DESCRIBING SUBSTANCES. . . .

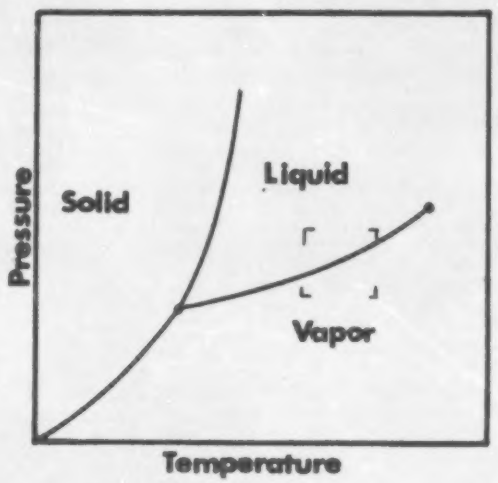
In part two of this module you found that the behavior of gases could be quite well described by the General Gas Law, expressed as a mathematical equation. This simple equation expressed the relation among the four parameters (P, V, N and T) that describes the state of a gas, and could be applied to any gas to predict its behavior when one of the parameters was changed. But how far can this Gas Law be used? If a gas condenses to form a liquid, for example, does the equation still apply, or does another one take over? Or is there a bigger, more grand equation, say a General Substance Law, which will simply express the relations between P, V, T and N over all their possible values, and which can be used to predict boiling points, melting points, volumes, pressures, and so on?

The answer is no. The behavior of substances is much too complicated for such a universal relation. No simple, single equation will do for all the phases of a substance what the General Gas Law does for the gas phase. Physicists have developed equations which describe many aspects of substances' behavior and they are currently hard at work extending their mathematical description. But in general, these equations and calculations are complicated and not of convenient utility to the worker in the field. Thus another way of describing the behavior of substances in required. In general this is simply a way of presenting the experimental data that has been taken on a substance in such a way that it can be easily used and various relationships seen. As in your experiments this very often means a graph.

. . . .USING GRAPHS:

The graphs which display the various phases of substances are quite logically called *phase diagrams*. Since graphs only have two axes, the phase diagram can only display the relation between two of the parameters at a time and many phase diagrams would be required to completely describe a substance. In general, however, one is only interested in certain pairs of relations and these lead to the two most common phase diagrams for substances, the P-T diagram and the P-V diagram.

In the pressure cooker, and thus for this module, it is the P-T diagram for water that is of interest. The illustration shows a schematic P-T diagram for an imaginary substance. As shown, pressure is on the vertical axis and temperature on the horizontal axis. The lines separate the graph into regions of solid, liquid and vapor. Each of these regions de-



A schematic diagram of a typical P-T diagram. In your experiment you will trace out a small part (shown boxed) of the liquid-vapor line for water.

fine all of the possible values of P and T for which the substances will be in that particular phase.

In this part of the module you will experimentally measure part of the phase diagram of water. You will see how the phase diagram can explain many common properties of water including evaporation, relative humidity, dew point, ice skating, the disappearance of snow without melting, and so on. You will then be able to interpret the phase diagram of other substances with which you are not so familiar.

YOUR EXPERIMENTS

TRACING OUT PART OF THE PHASE DIAGRAM. . . . . .

In your experiment for this part of the module you will operate the pressure cooker just as you would in the home. You will put some water in the bottom, bring it to a boil and then let the pressure rise until the weight begins to lift off to regulate the pressure. As the pressure increases you will measure the temperature of the steam with your thermistor and the pressure with the pressure gauge. If the temperature of the pressure cooker does not increase too rapidly, the liquid and the vapor will be in thermal equilibrium, that is the steam temperature will be the same as that of the liquid. Thus you will be measuring the change in boiling temperature with pressure.

In taking this data, you will be following along one of the lines of the phase diagram for water, the line separating the liquid and vapor regions. This line gives the pressure-temperature relation when the liquid is in equilibrium with its vapor. It will only be a small section of the diagram to be sure, but your experimental technique could be quite easily extended to cover the entire liquid-vapor line.

. . . . . .AND APPLYING YOUR RESULTS

As you will see later in this module, the data that you take has considerable importance and many practical applications. With your data you will be able to estimate changes in cooking time that result from differences in the altitude of where you do the cooking. You can determine how long you have to boil something to sterilize it, taking into consideration the pressure where you are. You will also see the considerable advantage of the pressure cooker for such purposes, and why the autoclave is an essential piece of hospital equipment. The reason behind the precautions in using these devices will be simply explained in terms of the phase diagram. And you can extend these analyses to other substances such as CO_2, anti-freeze, propane, freon, and so on.

PHASE DIAGRAMS AND THE WEATHER:

The amount of water vapor in the air and its behavior with respect to changes in temperature and pressure are of considerable importance in meteorology and weather prediction. Many things of interest such as fog, relative humidity, and dew point can be simply explained by looking at the phase diagram of water. Before you begin your main experiment on water we will show you some simple methods for measuring relative humidity and dew point using your thermistor. Later when you better understand the phase diagram you will see how it is that these methods give you the results and why these two quantities are of interest to weather men.

SOME SIMPLE EXPERIMENTS TO CHECK UP ON YOUR WEATHER MAN

THERE'S MORE TO WEATHER THAN TEMPERATURE AND PRESSURE. . . . .

In previous experiments you have been asked to find the temperature and pressure of the air in the lab using a thermometer and a barometer. Of course, these measurements could just as easily be made outside as inside, and hence it is easy to compare your results with those given in weather reports. There are a number of other things that are of interest about the weather, however. For instance, in the following experiments you will measure two quantities related to the amount of water in the air - the *relative humidity*, and the *dew point*.

EVAPORATIONAL COOLING. . . .

A wet thermometer gives a temperature different from a dry one, due to the evaporation of the wetting substance. To feel this difference, dip one finger in water and blow on that finger and the one next to it. Note that the wet one feels cooler. Now try rubbing alcohol on your finger. Compare this to both a dry finger, and one dipped in water. The difference in feeling can be made more precise by using a more accurate thermometer than your finger.

Evaporational cooling can be easily demonstrated by blowing on two fingers one of which is wet. The wet finger feels significantly colder then the dry.

A quantitative measure of evaporational cooling can be made by waving a wet thermistor in the air. The resulting equilibrium temperature of the thermistor is related to the relative humidity.

To make the measurement, first measure the room temperature using your thermistor as in previous experiments. Record this in the table on the opposite page.

Now wrap the thermistor in a very small water-soaked piece of paper towel or cotton. Be sure the water you use is at room temperature. Wave the wrapped thermistor in the air fairly briskly until its temperature stops changing. Record this as T_w. Note that the evaporating water has cooled the thermistor considerably (if the day is a "dry" one), and by a specific amount. Now calculate the difference $\Delta T = T_d - T_w$ and record it.

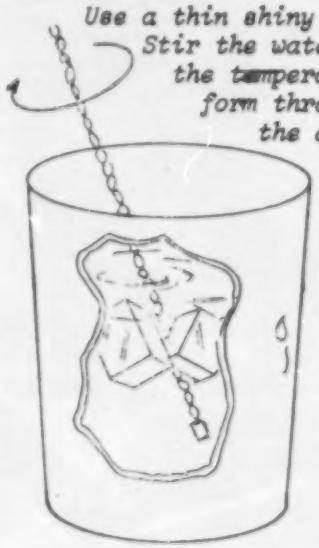
To see that the value of T_w depends on the fact that water was used, try "wetting" cotton or paper towel with rubbing alcohol, and find a "wet" temperature as before. For a different effect try some light oil, such as 3 in 1 oil. Since alcohol and oil both are practically absent from the air, it seems reasonable that T_w depends on the properties of the "wetting" substance, as well as how much of this substance is in the air.

RELATIVE HUMIDITY. . . . .

Later in this part we will explain exactly what "relative humidity" is. For the present, we will simply say it is a measure of how much water vapor there is in the air.

To convert your evaporational cooling temperatures refer to the chart on page . Note that the temperatures on the chart are in °F. Go down the column marked "Dry Bulb Reading" until you reach T_d, then go across the row until you reach the value under the "Dry-Wet Difference" equal to ΔT. This number is the relative humidity for the particular time and place you took the measurements.

Dry Thermistor Temperature:
T_d = _____
Wet Thermistor Temperature:
T_w = _____
Temperature Difference:
$\Delta T = T_d - T_w$ = _____
_____
Relative Humidity:
(from Chart page)
R.H.= _____%
Dew Point:_____



When the cup gets cool enough, the water vapor in the air condenses on it. The temperature at which this happens is called the dew point.

FINDING THE DEW POINT. . . . .

If the air in your room were suddenly cooled to a low enough temperature, the water vapor in the air would condense and the room would become foggy. The *temperature* at which this occurs is called the dew point, and it depends only on the amount of water vapor in the air.

A simple way to measure this temperature is to cool a thin metal cup with ice water while measuring the temperature of the water. When the dew point is reached, the water vapor in the air near the cup will condense onto the cup. To perform the experiment, put warm water and ice in a metal cup and stir the mixture constantly with the thermistor. This keeps the temperature throughout the cup uniform. Record the temperature exactly at the moment when water droplets appear to form on the cup. This temperature is called the *dew point temperature*. Record your value in the table.

A MEASUREMENT OF $T_{boiling}$ vs P:

GENERAL PROCEDURE AND PURPOSE. . . . . .

In this experiment you will operate the pressure cooker just as it would be operated in the home, except that you will measure the temperature and pressure inside the cooker as it heats up. In this way you will be able to determine how the temperature of boiling water depends on the pressure of the vapor above it. Also, you will be able to verify the pressure force relation (F = P x A) by noting the action of the pressure regulator.

THE STEPS

PREPARATION. . . .

* Make a data table like the one for part two of this module, with columns for resistance, pressure, and temperature. It is convenient to take data at one psi intervals starting at 0 and going up, so you can enter the pressure values in your table now.

* Weigh the pressure regulator and measure the diameter of the hole it fits on.

* Calculate the pressure at which you believe the regulator will lift off. Refer back to the sample calculation in part two.

Pressure Regulator:
m = _____

Hole:
d = _____

Calculated lift off Pressure:
$P_{cal.}$ = _____

Observed lift off Pressure:
$P_{abs.}$ = _____

Atmospheric Pressure:
$P_{atm.}$ = _____

* Attach the thermistor to the cover and check to see that it is off the bottom and operating properly. (Zero the meter and check room temperature and maybe one other point against your calibration graph).

* Find the atmospheric pressure in your lab and record it.

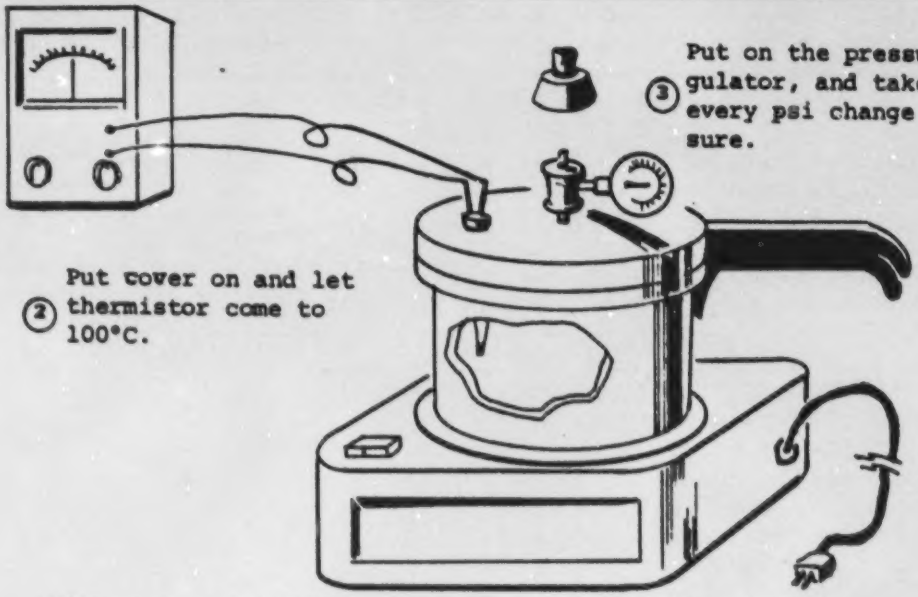
* Fill the pressure cooker about 1 cm. deep with water.

THE EXPERIMENT. . . . .

1. Place the pressure cooker on the hot plate with the cover off. Turn on the hot plate and let the water come to a boil. You will leave the pressure cooker on the hot plate throughout the experiment.

2. Put the cover on, but *don't put the pressure regulator on yet.*

THE PROCEDURE


Put on the pressure regulator, and take data every psi change in pressure.
3
Put cover on and let thermistor come to 100°C.
2
With cover off, bring 1 cm. deep H_2O to a boil in bottom of pressure cooker.
1
Take the cooker off the hot plate when P and T stop changing.
4

CAUTION: *Let the pressure return to near 0 before removing the pressure regulator.*

3. When the thermistor has reached the resistance corresponding to 100°C, put on the pressure regulator (*not the rubber stopper*) and begin taking data. Tap the pressure gauge frequently to make sure it is registering the correct value. Don't be concerned about the "hissing" of the top with the regulator on it; the seal is sufficiently good for this experiment.

   The regulator should ultimately start rocking back and forth, supported by the pressure of the escaping steam. Watch for a minute or two to verify that the equilibrium between pressure and gravity forces that was discussed in part two has been reached. (There should be no further P or T changes). Record the pressure at which the regulator released. How does this value compare with your calculations!

4. Remove the pressure cooker from the hot plate. Caution: *Let it cool or run it under cold water until the pressure approaches 0 psi before removing the pressure regulator.* As the thermistor shows, the steam gets quite hot, and since it would release all its latent heat if it condensed on your hands, this cooling procedure is important!

UNDERSTANDING YOUR EXPERIMENT....

Now that you have taken the data let us take a moment to think about the experiment to be sure we understand what the data means. At the start of the experiment the water inside the pressure cooker was boiling, steam was coming off and escaping through the hole at the top with no pressure rise. Then we put the regulator on so that the steam could not escape. The boiling, however, continued (perhaps you could hear it), converting more and more of the water into vapor. As the amount of vapor above the liquid increased the pressure rose and so did the temperature of the water. Thus the temperature at which boiling occurs depends on the pressure on the liquid. To understand why means understanding what we mean by boiling.

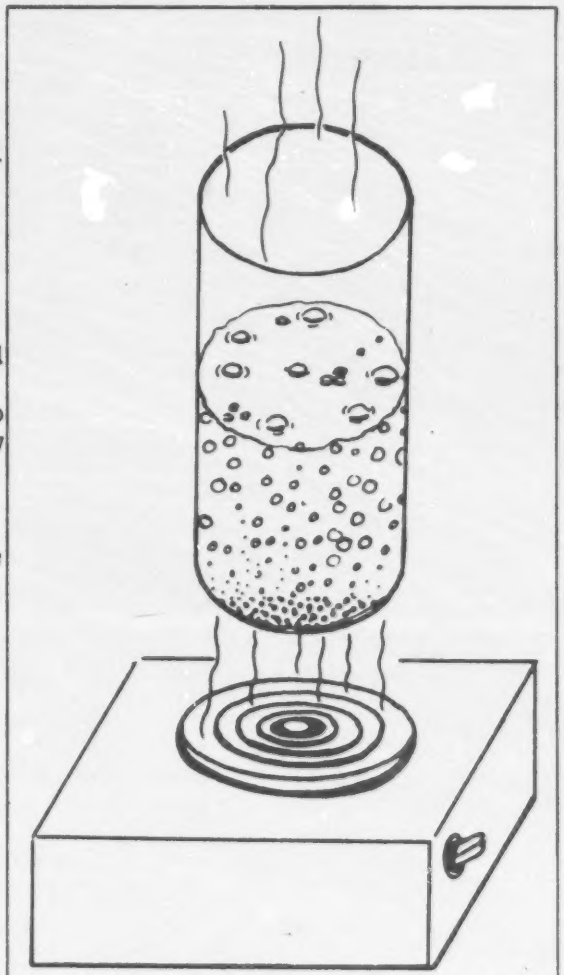
BOILING - One Type Of Vaporization:

Boiling is one example of the process of *vaporization*. (The other is evaporation which we will discuss later). Vaporization simply means the process of changing from a solid or liquid into a vapor. In the case of boiling this process is a violent one. Vapor bubbles form deep in the liquid (wherever the heat is applied) and float to the surface where the vapor, and the energy that went to create it, is released. As you saw in part one, when boiling begins, all of the input energy goes into creating these vapor bubbles and the temperature of the liquid *does not change*.

BOILING TEMPERATURE - Depends On Total Pressure:

In this experiment the boiling temperature did rise however, as the pressure in the cooker increased. Why did this happen? The explanation lies in the formation of the bubbles. Bubbles are continually trying to form in the liquid whenever heat is applied. However, if the pressure in the liquid is too great, the bubbles will collapse and never make it to the surface. The energy that went into trying to form the bubble is then dissipated in the liquid and its temperature rises.

When 212°F is reached however (and the pressure is one atmosphere) the bubbles are suddenly able to maintain themselves and boiling begins. Thus the abil-



Boiling is a process of vaporization in which all of the input energy goes to forming vapor bubbles. The bubbles float to the surface releasing the vapor and thus the energy that went to create them. The temperature at which the bubbles can form, that is the boiling temperature, is determined by the total pressure on the fluid. The greater the pressure, the higher must be the temperature for bubble formation.

ity to form bubbles is dependent on both temperature *and pressure*. If when boiling has started we suddenly increase the pressure the bubbles will collapse and boiling will stop. The heat will then go into raising the liquid temperature until it reaches a point where the bubbles can again maintain themselves. Boiling will then start at this new, higher temperature and pressure. Thus we see that the ability of a liquid to boil depends on both its temperature, as well as the *total* pressure on it.

In this experiment, the increase in pressure was produced by the vapor itself. An equilibrium situation was created in which the boiling temperature increased to keep pace with the increase in pressure. Your data then represents *the equilibrium conditions of temperature and pressure for boiling to occur in water.*

PLOTTING YOUR DATA

Your data analysis will be to plot a graph of boiling temperature, as measured with your thermistor, against the pressure in the cooker. The resulting curve will show exactly how the temperature at which water will boil depends on its total pressure, over a limited range of values. The procedure below may serve as a guide in preparing your graph:

1. COMPLETE THE DATA TABLE-As in the last experiment you must first convert your thermistor resistance values to temperatures using your calibration graph. Also your gauge pressure readings must be converted into absolute values by adding the value of atmospheric pressure for the day. Since your pressure data was taken in psi it is probably most convenient to use these units.

2. PLOT THE DATA-For your graph you should use linear graph paper, plotting pressure on the vertical axis and temperature on the horizontal axis. Let the pressure axis start at 0 psi absolute and the temperature axis at around 80°C. Choose convenient divisions for your scales but let your data fill as much of the graph paper as possible.

3. DRAW A SMOOTH CURVE-Draw the best smooth curve you can through your data points. As usual be sure that as many points lie above the curve as below. Since you are quite confident of the first point (P = 14.7 psi, T = 100°C) make your curve pass through that point.

4. COMPARE YOUR DATA-The data shown opposite is the published data on the boiling point of water over the range of your experiment. Plot a number of these data points on your graph. How well does your curve compare? Can you explain the irregularities in your data in terms of the experiment?

BOILING TEMPERATURE (°C)	PRESSURE (psi)
95	12.24
100	14.70
105	17.52
110	20.78
115	24.52
120	28.80
125	33.66
130	39.18
135	45.39

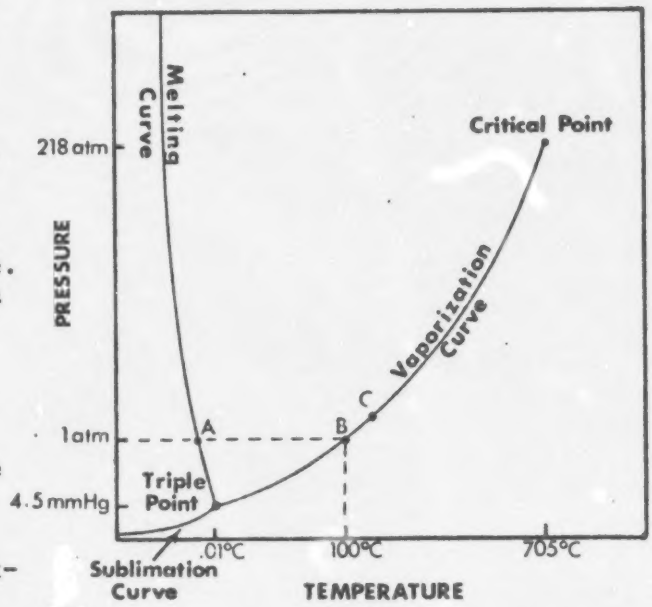
Table of published data on the variation of boiling temperature of water with pressure. For data over a wider range of temperature and for intermediate values see The Handbook Of Chemistry And Physics under "Vapor Pressure of Water".

YOUR GRAPH - A Small Part Of The Phase Diagram:

As we mentioned in the introduction, it is not possible to describe the behavior of substances undergoing phase changes by a single equation as we were able to do with the gas phase. Instead we generally resort to what is called a *phase diagram.* There are a number of different types of phase diagrams that display different properties of the substance, but we will be concerned here with only the simplest of these, the P-T diagram. The P-T diagram of water is shown below. *Note that the diagram is only schematic.* That is it shows the general phase diagram shape but the P-T scales are not linear. This is done so that all of the interesting points of the graph can be displayed at once. Generally only some small section of the graph will be of interest at a given time. Several of these are discussed on the opposite page and displayed to the proper scale.

INTERPRETING THE PHASE DIAGRAM. . . .

Regions: Take a moment to look at the phase diagram. What can it tell you? First, what are the regions? The lines divide the graph into three regions marked solid, liquid, and vapor. Each region specifies all of the conditions of pressure and temperature under which the substance will be in that phase. For example, if the pressure on water was 3 atm. and its temperature was 60°, that point lies in the liquid region and the water would be a liquid. But if at 3 atm. the temperature was only -20°C it would be a solid, ice...and so on.



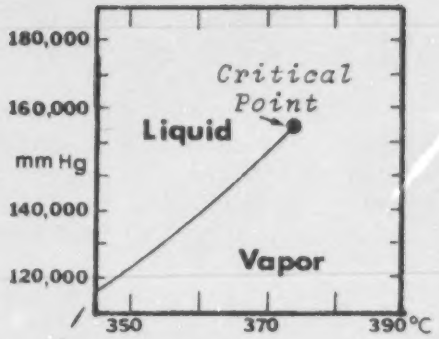
Schematic Phase Diagram Of Water.

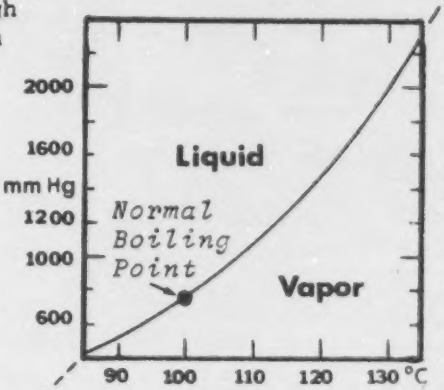
Phase Changes: Now what about phase changes. In your first experiment in this module you started with ice at atmospheric pressure and 0°C, point A in the diagram. You slowly increased the temperature while maintaining the pressure at one atmosphere. Your experiment followed the horizontal line through the solid-liquid phase change, through all of the liquid region and to the liquid-vapor phase change at point B. If you had performed the same experiment at a pressure of 1 mm Hg you would have completely missed the liquid phase and gone directly from solid to vapor (ice to steam), a process called *sublimation.* In this experiment you started at the boiling point of water (point B) and followed along the vaporization curve, letting the boiling pressure and temperature increase, to some higher boiling point (point C).

Another interesting path through the phase diagram is a vertical one. Suppose you have ice at 1 atm. and 0°C and you increase the pressure while maintaining the same temperature. What happens? According to the diagram, the ice melts. This phenomenon is responsible for our ability to ice skate. The weight of your body on the skates causes the runners to exert a large pressure on the ice. The ice under the runners melts and you skate smoothly along on a path of water.

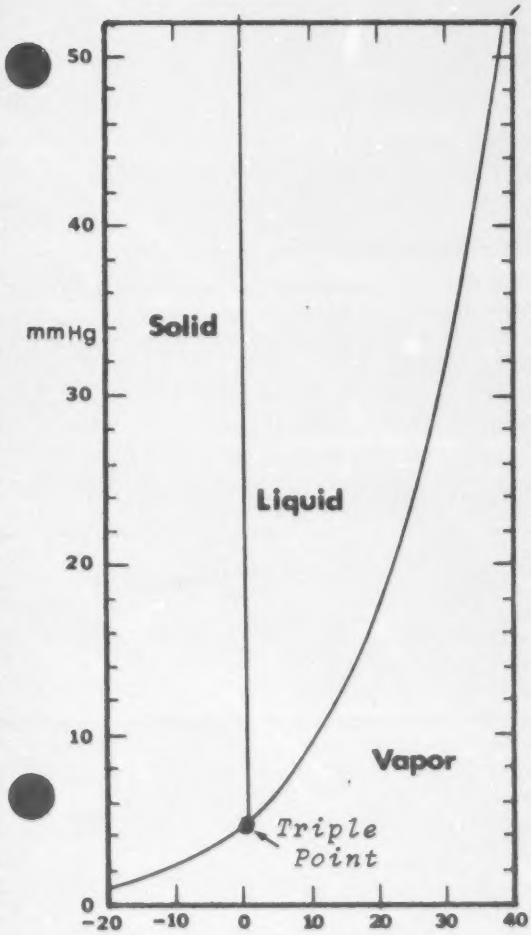
THE CRITICAL POINT. . . . .

Above the critical point there is *no distinction between the liquid and the vapor phase.* For water this occurs at an extremely high pressure (218 atm) and temperature (374°C). If you were heating water in a strong glass container the usual line that separates the liquid on the bottom from the vapor above would suddenly disappear when you passed through the critical point. There would be no separate liquid and vapor phases. All the water would be a gas of very high density. No matter how high you raise the pressure, you cannot liquify water above the critical temperature.





Three sections of the phase diagram of water shown to proper scale.



YOUR EXPERIMENT. . . . .

In your experiment with the pressure cooker, you started at the normal boiling conditions for water (760 mm Hg and 100°C) and moved up the *saturated vapor pressure curve.* For all points on this curve, the liquid and the vapor are in equilibrium; that is, the relative amounts of vapor and liquid are specified and do not change. The words "saturated vapor pressure" simply mean that the amount of water that will convert to the vapor phase (as measured by its pressure) has reached its maximum value, or is saturated. As you increased the temperature of the pressure cooker you continued to increase the amount that could become vapor; that is, raised its saturated vapor pressure.

EVAPORATION, HUMIDITY AND DEW POINT. . . . .

Evaporation, humidity, and dew point are determined by the section of the saturated vapor pressure curve that lies at normal outdoor temperature. From -15°C to 40°C (5°F to 104°F) the saturated vapor pressure varies between about 1 mm Hg and 55 mm Hg, a small fraction of one atmosphere. These pressures represent the maximum amount of water vapor that the air can hold at that temperature. In the following pages we will explore this region in more detail in order to explain why evaporation occurs and what your values for relative humidity and dew point represent.

THE TRIPLE POINT. . . . .

The triple point is the unique pressure and temperature in the phase diagram (4.58 mm Hg and .01°C) at which all three phase boundries come together. Under these conditions *all three phases are in equilibrium together,* that is the water will be boiling with ice floating in it and steam existing above it.

Since the triple point occurs at a unique temperature, this temperature (.01°C) is often used for calibrating thermometers.

EVAPORATION - Another Type Of Vaporization:

Boiling is not the only type of vaporization. When wet clothes are left out on a line, they dry; after a rain, the puddles disappear; when you blew on your wet fingers they quickly dried off. In each of these cases, a liquid has converted to a vapor without boiling.

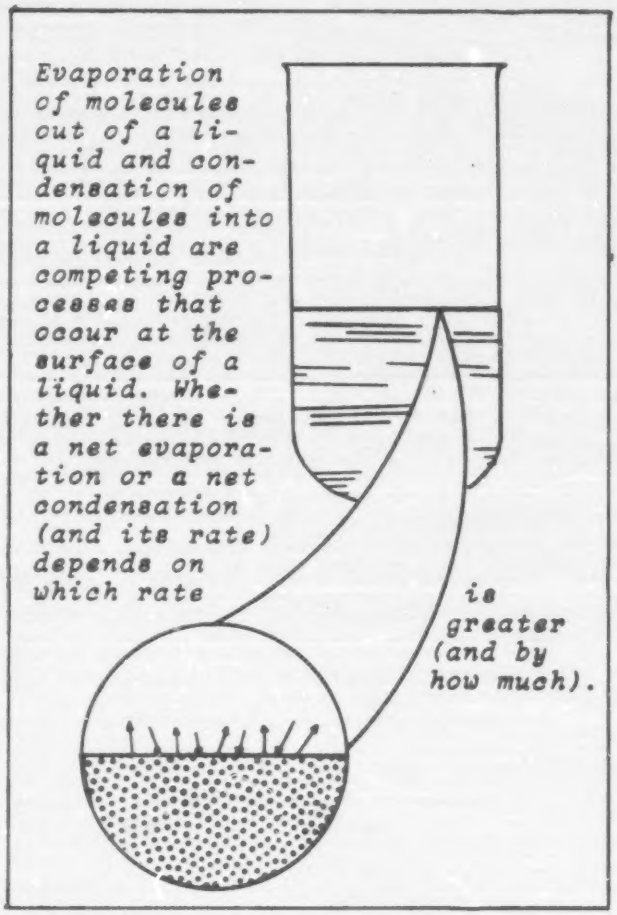
Evaporation is the process by which liquids can vaporize without boiling. It occurs at the *surface* of liquids rather than deep inside as in boiling. Every so often a molecule will have enough energy to break through the surface and escape. This in fact happens many times per second so that under the right conditions, evaporation can rapidly vaporize a liquid away.

And what are these conditions? As you would expect, one of them is temperature. The higher the temperature of a liquid, the more energy its molecules have and the greater the chance that a molecule will have enough energy to escape. Thus, *the greater the temperature the greater the rate of evaporation.*

CONDENSATION -

The Reverse Process:

However, evaporation is somewhat of a "two way street". That is if there are vapor molecules in the air above the liquid, many of them will be striking the surface of the liquid and will stick there. This process is called *condensation*. There will thus be a certain rate of condensation of vapor into the liquid at the same time there is evaporation out of it. The condensation rate will depend on the amount of water vapor in the air and this amount is specified by the *partial pressure* of the vapor in the air. As we will see, the partial pressure of water vapor in the air is closely related to the relative humidity.


Evaporation of molecules out of a liquid and condensation of molecules into a liquid are competing processes that occur at the surface of a liquid. Whether there is a net evaporation or a net condensation (and its rate) depends on which rate
is greater (and by how much).

EVAPORATION OR CONDENSATION? - Determined By The Partial Pressure:

When two or more gases are mixed together in the same space, the pressure of one alone is called its *partial pressure*. For example if air is composed of 20% oxygen, 78% nitrogen and 2% other gases, (and atmospheric pressure is 760 mm Hg), then the partial pressures will be:

Partial pressure of oxygen:	P = .20 x 760 mm Hg	= 152 mm Hg
Partial pressure of nitrogen:	P = .28 x 760 mm Hg	= 593 mm Hg
Partial pressure of other gases:	P = .02 x 760 mm Hg	= 15 mm Hg
Total Pressure:	P	= 760 mm Hg

When there is water vapor in the air, the amount of water can be specified by *its* partial pressure. And the *partial pressure of water vapor will determine the rate of condensation.*

In summary, then, the temperature of the water determines the rate of evaporation and the partial pressure of the vapor above the water determines the rate of condensation. Whether there will be a *net* evaporation or a *net* condensation will depend on which rate is greater; and the *net rate* will depend on how much greater. Relative humidity and dew point give a way of specifying these values.

THE DEW POINT

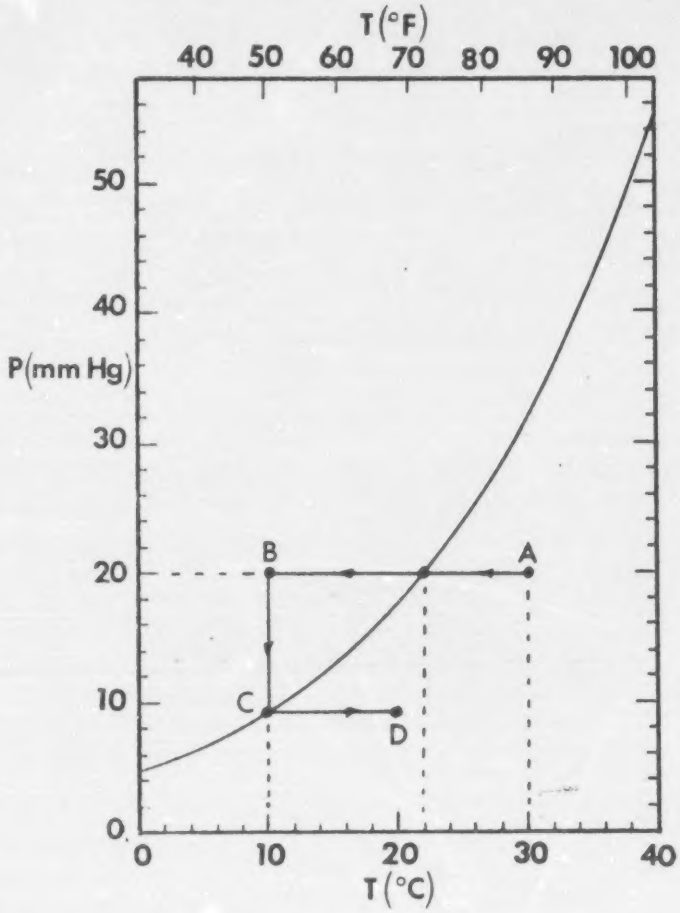
The *dew point* is that temperature at which the rates of evaporation and condensation are exactly equal. For temperatures above the dew point, the evaporation rate is greater and there is a net evaporation. When the temperature falls below the dew point, however, the condensation rate exceeds the evaporation rate and there is condensation. If the condensation occurs on objects (which often happens early in the morning) it is called *dew*. If the condensation is on particles in the air it is called fog or clouds. Earlier in the module you measured the dew point by placing a cold cup out in the air. As the cup cooled down, the rate of evaporation of the water steadily decreased. When the temperature passed through the dew point, the evaporation rate fell below the condensation rate and condensation droplets began to form on the side of the cup.

The dew point temperature, of course, varies with the amount of water vapor in the air. The greater the amount of water vapor, the higher the dew point. Thus, a high dew point generally means a humid day, a low dew point a dry day. By referring to the phase diagram of water, one can simply determine the amount of water vapor, as specified by its partial pressure, if one knows the dew point.

AN EXAMPLE:

The diagram below is the portion of the phase diagram relevant to the dew point. The curve is the liquid-vapor boundary and is the curve of temperature and pressure for which the liquid and vapor are in equilibrium, that is for which there is no net evaporation or condensation. The dew point temperature is represented by a point on this curve. The pressure scale represents the *partial pressure* of water vapor in the air.

For example, suppose the partial pressure of water in the air is 20 mm Hg. Then the dew point is 22°C. If the temperature outside is 30°C (point A) then water will evaporate. However, if the temperature drops to 10°C (point B) then condensation will begin. Condensation will continue until the partial pressure of water vapor in the air drops to that determined by the vapor pressure curve and the new temperature (point C). If the temperature warms up again (point D) evaporation begins and will continue until the partial pressure builds up to that determined by the curve at this temperature.



The dew point is of great practical importance for aviation and marine weather forecasting. For example; suppose the temperature at a certain airport is now 72°F and the dew point is 70°F. If the temperature falls only 2°F, fog could form and produce dangerously low visibility.

RELATIVE HUMIDITY

Relative humidity is another measure of the amount of water vapor in the air. Usually it is defined as the percentage ratio:

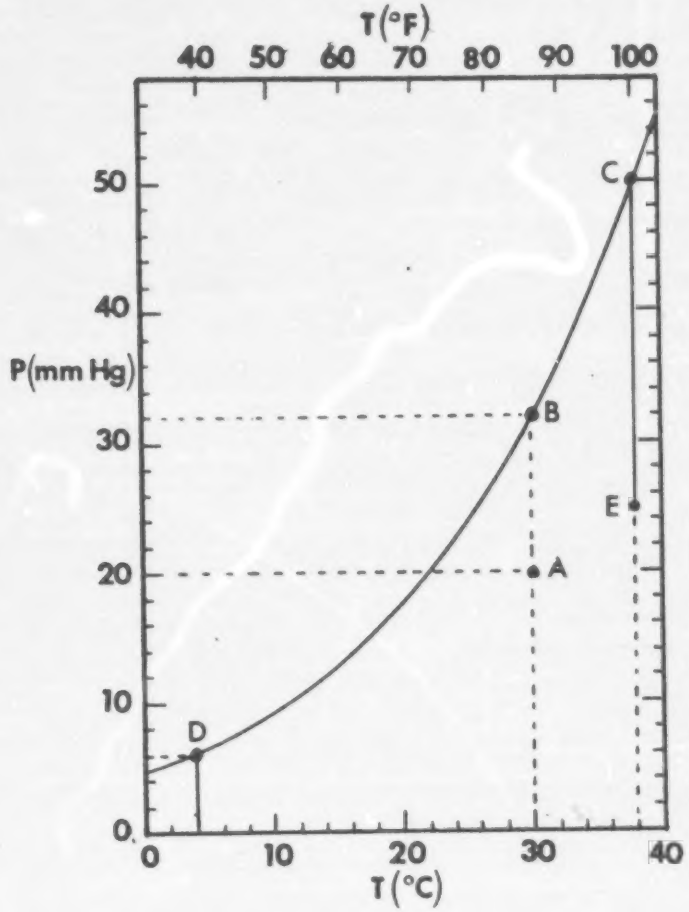
$$RH = \frac{\text{actual amount of moisture in air}}{\text{maximum amount of moisture air can hold}}$$

Thus, it represents the percent of water saturation of the air. If the air is quite unsaturated (a low relative humidity) then there will be a high net evaporation rate and things will dry out quickly.

On the other hand, if the water is highly saturated (a relative humidity near 100%) the temperature is close to the dew point and the evaporation and condensation rates are nearly equal. Under these conditions, evaporation is low and things take a long time to dry.

FOR EXAMPLE:

As with dew point, the phase diagram of water can be used to explain relative humidity. In terms of the phase diagram the relative humidity is defined by the percentage ratio: $RH = \frac{\text{actual partial pressure}}{\text{saturated partial pressure}}$



For example suppose the temperature is 30°C and the actual partial pressure of water in the air is 20 mm Hg (point A). The partial pressure required for saturation at that temperature is 32 mm Hg (point B). Thus the relative humidity is:

$$RH = \frac{20 \text{ mm Hg}}{32 \text{ mm Hg}} \times 100\% = 62.5\%$$

If the temperature changes, of course, so does the relative humidity.

EVAPORATION RATE & VAPOR PRESSURE:

The rate of evaporation, as we have seen, is dependent on the relative humidity, a high relative humidity means a low evaporation rate and vice versa. *However*, it is equally dependent on the absolute value of the saturated vapor pressure at that temperature. In fact what really determines the evaporation rate is the amount of partial pressure required to bring the air to saturation at that temperature. For example, at 38°C the saturated vapor pressure is 50 mm Hg (point C) while at 4°C it is only 6 mm Hg (point D). If the relative humidity at the lower temperature is 0% then evaporation must make up only 6 mm Hg partial pressure to become saturated and the net evaporation rate will be low. If at the higher temperature, on the other hand, the relative humidity is 50%, then the partial pressure is 25 mm Hg (point E) but evaporation must make up another 25 mm Hg partial pressure to become saturated. Evaporation will be much greater under these conditions than at the lower temperature.

Thus, it is the difference between the actual partial pressure and the saturated partial pressure that determines the evaporaiton rate. *The evaporation rate is directly proportional to this partial pressure difference.* This explains why things will dry much more quickly on a hot dry day than on a cold dry day.

PSYCHROMETRIC TABLE FOR RELATIVE HUMIDITY

The table below gives the relative humidity as determined from wet and dry bulb readings of temperature. It is called a psychrometric table after this technique for measuring the relative humidity. That is, a psychrometer is the meteorological term for a wet and dry bulb hygrometer, and a hygrometer is simply any instrument used for measuring the amount of moisture in the air.

EXAMPLE

If dry bulb reading is 68° and wet bulb reading is 58° the relative humidity will be found in the 10° difference column opposite 68°. In this case 55%.

DRY BULB READING	DRY-WET DIFFERENCE 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
30	89	78	68	57	47	37	27	17	8																					
32	90	79	69	60	50	41	31	22	13	4																				
34	90	81	72	62	53	44	35	27	18	9	1																			
36	91	82	73	65	56	48	39	31	23	14	6																			
38	91	83	75	67	59	51	43	35	27	19	12	4																		
40	92	84	76	68	61	53	46	38	31	23	16	9	2																	
42	92	85	77	70	62	55	48	41	34	28	21	14	7																	
44	93	85	78	72	64	57	51	44	37	31	24	18	12	5																
46	93	86	79	72	65	59	53	46	40	34	28	22	16	10	4															
48	93	87	80	73	67	60	54	48	42	36	31	25	19	14	8	3														
50	93	87	81	74	68	62	56	50	44	39	33	28	22	17	12	7	2													
52	94	88	81	75	69	63	58	52	46	41	36	30	25	20	15	10	6													
54	94	88	82	76	70	65	59	54	48	43	38	33	28	23	18	14	9	5												
56	94	88	82	77	71	66	61	55	50	45	40	35	31	26	21	17	12	8	4											
58	94	89	83	77	72	67	62	57	52	47	42	38	33	28	24	20	15	11	7	3										
60	94	89	84	78	73	68	63	58	53	49	44	40	35	31	27	22	18	14	10	6	2									
62	94	89	84	79	74	69	64	60	55	50	46	41	37	33	29	25	21	17	13	9	6	2								
64	95	90	85	79	75	70	66	61	56	52	48	43	39	35	31	27	23	20	16	12	9	5	2							
66	95	90	85	80	76	71	66	62	58	53	49	45	41	37	33	29	26	22	18	15	11	8	5	1						
68	95	90	85	81	76	72	67	63	59	55	51	47	43	39	35	31	28	24	21	17	14	11	8	4	1					
70	95	90	86	81	77	72	68	64	60	56	52	48	44	40	37	33	30	26	23	20	17	13	10	7	4	1				
72	95	91	86	82	78	73	69	65	61	57	53	49	46	42	39	35	32	29	25	22	19	16	13	10	7	4	1			
74	95	91	86	82	78	74	70	66	62	58	54	51	47	44	40	37	34	30	27	24	21	18	15	12	9	7	4	1		
76	96	91	87	83	78	74	70	67	63	59	55	52	48	45	42	38	35	32	29	26	23	20	17	14	12	9	6	4	1	
78	96	91	87	83	79	75	71	67	64	60	57	53	50	46	43	40	37	34	31	28	25	22	19	16	14	11	9	6	4	1
80	96	91	87	83	79	76	72	68	64	61	57	54	51	47	44	41	38	35	32	29	27	24	21	18	16	13	11	8	6	4
82	96	92	88	84	80	76	72	69	65	62	58	55	52	49	46	43	40	37	34	31	28	25	23	20	18	15	13	10	8	6
84	96	92	88	84	80	77	73	70	66	63	59	56	53	50	47	44	41	38	35	32	30	27	25	22	20	17	15	12	10	8
86	96	92	88	85	81	77	74	70	67	63	60	57	54	51	48	45	42	39	37	34	31	29	26	24	21	19	17	14	12	10
88	96	92	88	85	81	78	74	71	67	64	62	58	55	52	49	46	43	41	38	35	33	30	28	25	23	21	18	16	14	12
90	96	92	89	85	81	78	75	71	68	65	62	59	56	53	50	47	44	42	39	37	34	32	29	27	24	22	20	18	16	14
92	96	92	89	85	82	78	75	72	69	65	62	59	57	54	51	48	45	43	40	38	35	33	30	28	26	24	22	19	17	15
94	96	93	89	86	82	79	75	72	69	66	63	60	57	54	52	49	46	44	41	39	36	34	32	29	27	25	23	21	19	17
96	96	93	89	86	82	79	76	73	70	67	64	61	58	55	53	50	47	45	42	40	37	35	33	31	29	26	24	22	20	18
98	96	93	89	85	83	79	76	73	70	67	64	61	59	56	53	51	48	46	43	41	39	36	34	32	30	28	26	24	22	20
100	96	93	90	86	83	80	77	74	71	68	65	62	59	57	54	52	49	47	44	42	40	37	35	33	31	29	27	25	23	21
102	96	93	90	86	83	80	77	74	71	68	65	63	60	57	55	52	50	47	45	43	41	38	36	34	32	30	28	26	24	22
104	97	93	90	87	84	80	77	74	72	69	66	63	61	58	56	53	51	48	46	44	41	39	37	35	33	31	29	27	25	24
106	97	93	90	87	84	81	78	75	72	69	66	64	61	59	56	54	51	49	47	45	42	40	38	36	34	32	30	28	27	25
108	97	93	90	87	84	81	78	75	72	70	67	64	62	59	57	54	52	50	47	45	43	41	39	37	35	33	31	29	28	26

Review

SUMMARY

This part of the module is mainly concerned with changes between the liquid and vapor (or gas) phases. These phase changes depend on the temperature and also on the total pressure applied to the material.

In the experiment, you measured the boiling temperature of water in a pressure cooker, as the total pressure in the cooker changed. The boiling point was found to increase as the pressure increased.

Boiling is a special case of the general process of *vaporization*, which refers to the change from solid or liquid to gas. Boiling occurs when vapor bubbles form in the liquid and rise to the surface without collapsing.

Evaporation is a slow process of vaporization that goes on all the time at the surface of liquids. *Condensation* is the reverse of vaporization: it is the change from vapor to liquid or solid.

The number of vapor molecules present determines the *vapor pressure* in a region. If several gases are present, then the pressure resulting from each one is called its *partial pressure*. When a liquid and its vapor exist together, the molecules are entering and leaving the liquid at the same rate, the two phases are in *equilibrium*.

Phase changes are often represented by graphs of pressure against temperature called *phase diagrams*. The lines on the graph represent boundaries between the solid, liquid, and gas phases. The line between the liquid and the vapor phases is a boiling point curve. The line between the solid and liquid phases is a melting point curve.

In your data analysis, you drew a graph of boiling point against pressure: this was a small part of the phase diagram for water. When the vapor is in equilibrium with the liquid, then the space above the liquid is said to be *saturated*. Your graph is also a plot of the *saturated vapor pressure* of water.

For each substance there is a unique temperature and pressure at which all three phases can exist together in equilibrium. This is called the *triple point*, and the temperature value for water (.01°C) is often used for calibrating thermometers.

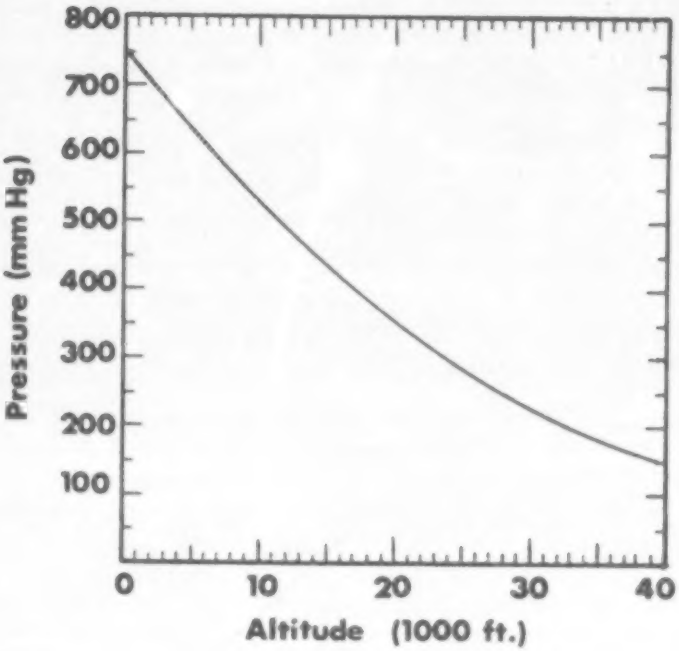
There is also a temperature, called the *critical* temperature, above which gases cannot be liquified. Propane is easily liquified under pressure, because its critical temperature is quite high.

Dew point is more accurately defined as the temperature at which the air becomes *saturated* with water vapor. That is, at the dew point the partial pressure of water vapor in the air equals the saturated vapor pressure for that temperature. Then the air is holding all the water vapor it can and condensation of the excess water begins.

Relative humidity is the ratio of the *actual* partial pressure of water vapor to the saturated vapor pressure. A measurement of temperature and dew point gives a good way to determine the relative humidity.

QUESTIONS

1. The temperature at your location is 47°F, there is no wind, and there is a dense fog. Estimate the dew point and the relative humidity.



Variation of atmospheric pressure with altitude above sea level.

2. The graph opposite shows how, on the average, atmospheric pressure varies with altitude. Of course, this graph is never perfectly accurate, because there are always changes in pressure due to weather and other causes. It does give you, however, a rough idea of the pressure variation. Notice that the pressure has fallen to about one-half the surface value at 18,000 feet.

   The summit of Mt. Everest is at a height of about 29,000 feet. What is atmospheric pressure at that height?

   Use the graph on page to find the boiling point of water on top of Mt. Everest.

3. The city of Denver, Colorado is at a height of about 5,000 ft. Find the average atmospheric pressure and boiling point of water in Denver.

4. You have just moved to Denver (see #3). You are cooking carrots in your pressure cooker. The regulator is set to hold the pressure at 15 psi gauge.

   a) What is the maximum *absolute* pressure in the cooker?

   b) What is the temperature in the cooker?

   c) The cooker instruction book says to cook carrots for 5 minutes; is that enough time?

5. Explain briefly, with reference to the phase diagram of water, why solid ice forms on the freezing coils of a refrigerator. Remember, there is no liquid present.

6. During World War II in England fog was sometimes temporarily removed from a small area (such as an aircraft landing strip) by burning large quantities of gasoline or fuel oil. Explain briefly.

7. When water is boiling in a very deep, glass container you can see the bubbles increase in size as they rise to the surface. Explain briefly.

8. At room temperature (20°C) could you liquify air by applying pressure? Explain briefly why or why not.

9. Why do clothes dry slowly when the relative humidity is high? Do clothes dry at all when hung out at temperatures below the freezing temperature of water?

10. In hot, dry places water is often kept in porous, clay pots. These let some water get to the outside of the pot, so it is always wet. Why do you think they do this?

PROBLEMS

1. On a certain very humid day, the air temperature is 86°F (30°C). Liquid moisture condenses on the outside of a tin cup when its temperature is 72°F (22°C).

   a) What is the dew point?

   b) Use the graph on page to find the *saturated* vapor pressure of water at the dew point temperature.

   c) What is the partial pressure of water vapor at the actual air temperature of 30°C?

   d) What *would be* the saturated water vapor pressure at 30°C *if* the air were saturated at that temperature?

   e) What is the relative humidity?

2. Thermistors are sometimes used as a kind of altimeter to measure altitude above sea level. Water is boiling in an open pan on an airplane with a non-pressurized cabin. A thermistor located just above the liquid surface indicates a temperature of 80°C. Use the graphs of this section, the graph of pressure vs altitude and any other information that you may need. Assume that atmospheric pressure has about its average value.

   a) What is the atmospheric pressure in mm Hg at the height of the airplane?

   b) What is the altitude of the airplane?

3. You are working at a biological field station at a height of 20,000 feet in the Andes Mountains of South America, You wish to adjust the laboratory autoclave to boil eggs in the same time that would be required at sea level.

   a) What temperature do you want inside the autoclave?

   b) What absolute pressure do you want inside the autoclave?

   c) At what gauge pressure do you want the regulator to release the steam?

"END OF DOCUMENT"